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**LABORATORY MANUAL AND PRINCIPLES  
OF CHEMISTRY**



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LABORATORY MANUAL

AND

PRINCIPLES OF CHEMISTRY

FOR BEGINNERS

BY

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## PREFACE.

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THIS course of experiments has been arranged for the elementary students of chemistry in this university. The number of quantitative experiments is greater than is commonly found in such laboratory manuals. The writer has found, however, that few of them are too difficult for the ordinary beginning student, and that careful workers can get very acceptable results in all cases. The beneficial results which must follow from giving students experiments which require some time, care, and thought, must be apparent to all teachers. Moreover, so many of the fundamental ideas of chemistry are based upon the quantitative relations of the elements, that such conceptions must be given to the student very early in his study of chemistry.

In Part II. the writer has attempted to give a clear and concise statement of some of the fundamental theories and principles of chemistry; here he has gone more into detail than is customary in books designed for elementary students. But as the elementary student very soon begins to use these theories and principles, it seems that he should very early have a fairly complete statement of them with the evidence upon which they are based.

All detailed treatment of chemical compounds and the chemistry of the individual elements has been purposely omitted, not from any lack of an appreciation of the value of such treatment, but because it would be out of place in a book principally intended as a *laboratory manual*.

At first the student should depend entirely upon his laboratory work, with such aid as he may get from his teacher for his chemical facts. After he has learned the value of careful observation, he may then be referred by his teacher to some good text-book, where he may extend his knowledge of the detail which is so indispensable to a thorough knowledge of chemistry, and which is best learned in the laboratory supplemented by careful reading in well-chosen text-books.

G. M. R.

STANFORD UNIVERSITY, CALIFORNIA,  
March 9, 1894.

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## NOTE FOR THE STUDENT.

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The object of this course in laboratory work is three-fold : —

(1) To acquire a knowledge of some of the fundamental principles upon which the science of chemistry is based.

(2) To acquire the habit of careful observation, and of deducing conclusions from such observations.

(3) To learn laboratory methods, to get habits of neatness and dispatch in setting up apparatus and conducting experiments.

To accomplish this it is indispensable that the work be done with great care. Keep your desk and apparatus clean and neat at all times ; no one who neglects this can do successful work in chemistry.

An experiment properly and intelligently performed is of value, but an experiment performed mechanically and without thought, or performed improperly and left without farther effort, does positive harm.

Time spent in *repeating* experiments until they are thoroughly understood, and until the facts they are intended to give are obtained, is well spent. Never leave an experiment until you fully understand it, and are satisfied with the results you have obtained.

It is not necessary to do *all* the experiments in the course. The quality of your work is much more important than the quantity.

Keep neat and careful notes of your work in a suitable blank-book. Begin the entry for each day with the date.

Do not crowd your notes ; leave room for remarks, additions, and corrections.

Refer to the experiment by number, and state concisely, what you did ; what you observed ; what you learned.

Give all the data furnished by the experiment, such as measurements, weights, etc., and make all the calculations in the note-book so that numerical errors as well as errors in method may be detected and corrected.

Provide yourself with a working apron to protect your clothing.

Always have a towel available.

Burns and acid wounds after washing out with cold water should be treated with a paste of sodium bicarbonate and water.

In case of a fire, a wet towel thrown over the flame will generally be sufficient to extinguish it.

*Experiments marked with a \* should be performed by the teacher before the class, or at least with the aid of the teacher, as they require special care in manipulation or require special forms of apparatus.*

*Experiments marked with a † should be performed in the hood or draught closet, or out of doors, as they develop harmful or disagreeable vapors.*



# LABORATORY MANUAL OF CHEMISTRY FOR BEGINNERS.



## PART I.



### EXPERIMENTS.

**The effect of *pressure* upon the volume of a gas, the *temperature* remaining constant.**

1. Allow the tube prepared for the purpose<sup>1</sup> to lie horizontally upon the desk; the air confined between the mercury and the sealed end of the tube is then under the atmospheric pressure. Note carefully the height of the barometer, which is the measure of the atmospheric pressure. Also measure the volume of the confined air; this may be done by simply measuring the length of the air column. (This assumes that the calibre of the tube is constant; in very accurate work the tube would need to be calibrated.) Now increase the pressure upon the air by

---

<sup>1</sup> For this experiment a glass tube  $1\frac{1}{2}$  m. long and 3 to 4 mm. internal diameter is carefully cleaned and dried; it is then heated about 40 cm. from one end and bent at right angles. Fill the tube, with the exception of about 20 cm. at each end, with clean dry mercury, and seal the open end of the short arm of the tube by bringing it into the flame of a Bunsen burner, thus confining a volume of air between the sealed end of the tube and the mercury.

raising the long and open arm of the tube above the table, the short arm lying in its horizontal position on the table. Bring the tube to rest at any desired position, measure the new volume of air, and measure the new pressure. Finally, bring the long arm to the vertical position and measure the pressure and volume. Make five or six measurements of pressures and corresponding volumes, one of which should be a pressure of exactly two atmospheres.

In each case multiply the number representing the pressure by the number representing the volume, and compare their products with one another. From this comparison what conclusions can be drawn in regard to the relation between the variable quantities, volume, and pressure?

Calculate what volume a liter of air measured at a pressure of 700 mm. would occupy under a pressure of 760 mm.

Deduce a general expression for the volume  $V$ , at pressure  $P$ , when the volume  $v$  and pressure  $p$  are known.

All gases behave toward changes in pressure just as air does. This law was discovered by Robert Boyle, in 1662, and is known as *Boyle's Law*. Edme Mariotte also discovered the same law in 1676; it is sometimes called *Mariotte's Law*.

**The effect of temperature upon the volume of a gas, the pressure remaining constant.**

2. Place the tube and thermometer arranged for the purpose,<sup>1</sup> in a cylinder of cold water (ice water if available).

---

<sup>1</sup> A simple apparatus for this experiment can be made by taking a clean dry glass tube 50 cm. long and 3 mm. internal diameter; put into the tube enough clean dry mercury to make a column 2 cm. long; allow this to run into the middle of the tube, and then seal one end. To this tube fasten a thermometer by means of a piece of wire so that

so that the whole air volume shall be immersed; allow to stand until the air volume suffers no farther change. Read the volume and temperature. Transfer the tube and thermometer to another cylinder containing warm water ( $40^{\circ}$ – $50^{\circ}$ ); after the air volume suffers no farther change, read the volume and temperature again. Finally, transfer the tube and thermometer to a large flask containing about 200 cc. of boiling water, from the neck of which steam is freely escaping; after the air volume has become constant, read the volume and thermometer. In reading the volume in each case, the tube should be held perfectly vertical. Why?

Is the expansion of air for each degree of temperature the same between the first and second reading, as between the second and third? In other words, is the coefficient of expansion of air constant between the extremes of temperature reached in the experiment?

Taking the volume at the lowest temperature ( $0^{\circ}$  if ice water is used) as unity, what fraction of this is the increase in volume due to a rise of one degree in temperature?

The volumes of all gases are affected by changes in temperature just as the volume of air is; they have the same coefficient of expansion that air has. This law was first discovered by Charles, and later by Dalton and Gay-Lussac, independently. It is referred to by different writers as the *Law of Charles*, or the *Law of Dalton*, or the *Law of Gay-Lussac*.

---

the bulb of the thermometer comes about opposite the middle of the air confined by the mercury in the sealed tube. This tube may be graduated, or the graduations on the thermometer may be used, and these readings finally converted into volume by measuring the length of the corresponding volumes with a metric rule.

**WATER.**

3. Write out briefly the characteristic properties of water from what you already know of it. At what temperature does it freeze? At what temperature does it boil? At what temperature does 1 cc. weigh the most? What is that weight?

4. Obtain a flask of about 750 cc. capacity, fill it  $\frac{1}{2}$  to  $\frac{3}{4}$  full of water from the hydrant, and connect with a *clean* condenser as shown in Fig. 1.



FIG. 1.

Boil the water, and collect in a clean beaker about 200 cc. of the water which distils over. Compare this distilled water with the hydrant water as to taste, odor, color, and effect upon red or blue litmus paper.

Evaporate a little of the distilled water in a porcelain dish; does it leave a residue?

Does the hydrant water leave a residue when evaporated?

Would sea water be rendered fit to drink by distillation?

5. Does the pressure under which water boils affect the temperature at which it boils? The effect of pressure upon

the boiling point of water may be studied as follows : Take a strong filtering flask, containing a little distilled water, and fit a rubber stopper carrying a thermometer into the neck of the flask ; set the flask upon an asbestos board or iron plate on a tripod. Connect the filtering flask with

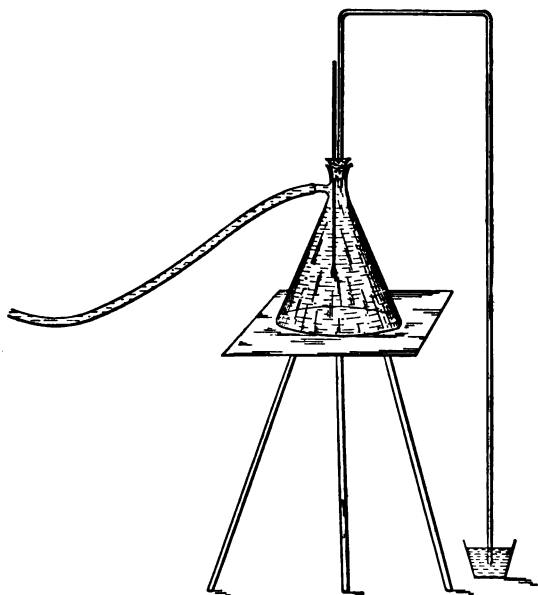


FIG. 2.

the filter pump, start the pump, and heat the asbestos board or iron plate. When the water in the flask begins to boil, read the temperature on the thermometer. The pressure in the flask under which the water is boiling can be measured as follows : Let the stopper which carries the thermometer have two holes ; into the second hole

fit a glass tube which will extend about 30 cm. above the stopper; then bending at right angles, extend out far enough to clear the tripod 20 or 25 cm., and bending at right angles again, extend down to the table; under the end of this open tube set a little dish of mercury so that the open end of the tube is below the surface of the mercury (see Fig. 2); as the pressure inside the flask diminishes, the atmospheric pressure will force the column of mercury up the tube until it compensates the diminished pressure within the flask; hence measuring the height of the mercury column and subtracting it from the height of the barometer observed at the time, the difference represents the pressure in millimeters of mercury under which the water in the flask is boiling.

Make four or five readings of the boiling point at different pressures; different pressures may be obtained by regulating the filter pump. Write the results in tabular form. Do equal diminutions of pressure produce equal changes in the boiling point?

6. Select some dry crystals of copper sulphate; put them in a test tube, heat the tube, and examine the liquid given off. For this purpose fit to the test tube a cork carrying a glass tube bent so that the liquid given off will collect in it and run out into a watch glass. What is the liquid?

7. Treat some alum crystals just as the copper sulphate crystals were treated; what liquid is given from these?

8. Treat some crystals of potassium bichromate in a similar manner; how do they differ from the copper sulphate and the alum crystals?

9. Carefully weigh two clean porcelain crucibles (No. 00); fill them  $\frac{1}{4}$  or  $\frac{1}{8}$  full of powdered copper sulphate crystals (select clean crystals to be powdered for this experiment)

and weigh again; then carefully heat the crucibles by allowing them to stand on a hot iron plate or asbestos board until the copper sulphate has lost its blue color; put the crucibles in a desiccator to cool, then weigh; after weighing, heat again for 20 minutes, allow to cool in the desiccator, and weigh. Repeat this process until the copper sulphate ceases to lose weight upon being heated. What per cent has the copper sulphate lost in weight in each case? Do the two samples lose the same per cent of their weight? To what is the loss due? See note in next experiment.

Save one crucible of the heated copper sulphate for use in Exp. 12.

**10.** Repeat previous experiment, using powdered alum crystals instead of copper sulphate. Do the two samples of alum lose the same percentage of their weight? Do they lose the same percentage of their weight that copper sulphate does?

N.B. To save time, the student should do this and the previous experiment together, heating and cooling all the crucibles at the same time, taking care, however, that the crucibles are marked so that they shall not become confused.

**11.** Heat about 15 grams of copper sulphate crystals in a small porcelain dish until they are white; after cooling add about 100 cc. of distilled water. What change do you observe? Finally, warm the water with occasional stirring until it is near boiling; when the bulk of the liquid has been reduced to about 20 cc., set the dish away in the desk and allow to stand several hours. A day or two will do no harm. Finally, examine the contents of the dish. What has taken place? Explain it.

**12.** Weigh accurately three porcelain crucibles (No. 00). Fill one  $\frac{1}{4}$  full of copper sulphate crystals; put in the second a similar amount of clear sodium sulphate crystals

(Glauber's salt) ; fill the third about  $\frac{1}{4}$  full of calcium chloride ; weigh each again ; take one of the crucibles of heated copper sulphate used in Exp. 9, the weight of which is known. Place all four of these away safely in your desk properly labelled, and loosely covered to keep out dust, and allow to remain one week. At the end of this time examine them and describe the changes that have taken place in each. What changes in weight have they undergone? Explain these changes.

In general, crystals which behave as the crystals of copper sulphate do are said to be *stable* ; those which behave as the crystals of sodium sulphate do are said to *effloresce*, or to be *efflorescent* ; and those which behave like calcium chloride crystals are said to *deliquesce*, or to be *deliquescent*. Write a definition for the terms *efflorescent* and *deliquescent*.

13. Take four small beakers ; put 20 cc. of water in each. Place in one 10 grams powdered potassium bichromate ; in another, 10 grams powdered common salt ; in the third, 10 grams powdered sodium sulphate ; in the fourth, 10 grams powdered alum : if, upon standing, all of the substance in any case is dissolved, add a little more.

Label each of these, and set them aside for an hour, occasionally shaking them ; in the mean time, procure four small beakers (about  $6 \times 4$  cm.), carefully clean and dry, label them, and weigh accurately. Note the temperature of the water in the above solutions ; then measure out carefully with a pipette 5 cc. of each solution into the different weighed beakers, taking the precaution to rinse out the pipette after each measurement ; weigh the beakers again with their contents. (Find the specific gravity of these solutions.) Then place the beakers on a warm plate to evaporate the water ; after they are completely dry, weigh again, taking care during the evaporation not to lose any



of the salts. From these data calculate how much of each salt is soluble in 100 parts of water. Write the results together in tabulated form. Are the salts left in the beaker the same as those dissolved in the first place?

**14.** Procure five small beakers,<sup>1</sup> label each, and cover it with a small watch glass; weigh accurately each beaker with its watch glass.

To another beaker containing 50 to 70 cc. of boiling water add 45 grams of powdered potassium bichromate with continued stirring until all or nearly all the bichromate has dissolved; then remove the heat. Put a thermometer into the solution, and when the temperature falls to 80°, transfer to one of the weighed beakers 5 cc. of the solution by means of a 5 cc. pipette, covering the beaker closely with the watch glass. When the solution in the larger beaker is cooled down to 65°, measure out another 5 cc. into a second beaker, as before; put in the remaining beakers 5 cc. each of the solution at 50°, at 35°, and at 20°. After the beakers are cold, weigh them with their contents; finally, evaporate and weigh the residues. From these results calculate how much potassium bichromate is soluble in 100 parts of water at the different temperatures. Write the results in tabular form. Also express them in the form of a curve. (See Fig. 41.)

NOTE. In measuring hot saturated solutions from a pipette, the pipette should be kept hot by keeping it in a vessel of boiling water. Why?

**15.** Repeat the previous experiment, using common salt instead of potassium bichromate. Compare the results. How do the curves differ?

---

<sup>1</sup> Small 50 cc. flasks, each closed with a cork, can be used to advantage instead of beakers.

16. Dissolve 50 grams of sodium sulphate in 25 cc. of water in a small flask, by heating the contents of the flask until the liquid boils. Cover the mouth of the flask with a card or piece of glass, and allow to stand at perfect rest until it has reached the temperature of the room. Does it crystallize? Put a stirring rod into the solution and stir, leaving the rod in the solution. What effect has this? Such a solution is said to be *supersaturated*.

17. Put about 10 cc. of water in a test tube; find the temperature of the water by inserting a thermometer; add a small piece (not more than 3 grams) of potassium hydroxide. What effect does this have upon the temperature of the water?

18. Repeat the previous experiment, using instead of potassium hydroxide about 3 grams of ammonium sulphocyanate. What effect does this have upon the temperature of the water?

19. Test the solubility of the following liquids in water: Alcohol, ether, coal oil, glycerine, carbon bisulphide, in the following manner: In a clean test tube place 5 cc. of distilled water; pour into this 1 cc. of the liquid to be tested; shake several times and then allow to settle and observe the mixture.

Gases also dissolve in water, as we shall see farther on.

20. Mix intimately 10 grams of sodium carbonate and 10 grams of tartaric acid in a mortar. Are there any changes? Take a little of the mixture in a beaker and add water. What occurs? Does water act upon either sodium carbonate or tartaric acid alone as it does upon the mixture? Give an explanation of the behavior here.

21. Place 5 cc. of alcohol in a test tube and add 1 cc. of strong hydrochloric acid; shake well, and then add a piece of fused potassium carbonate.

Put 5 cc. of water in another tube and add 1 cc. of strong hydrochloric acid. Drop a small piece of potassium carbonate into the tube. Note the difference. Can you explain this?

**22.** Repeat the previous experiment, using a piece of marble instead of a piece of potassium carbonate. Explain the difference in action between the potassium carbonate and the marble.

**\*23.** Take a wide-mouthed bottle, capable of holding 100 cc. or more; fill it completely with pure water; fit in the neck a cork carrying a delivery tube of glass. Take two platinum wires about 8 cm. long, and fasten to each a piece of platinum foil about  $1 \times 3$  cm. Dip the foils into the water and allow the wires to extend out of the bottle either through the cork or at the side of the cork. (See Fig. 3.) Cork tightly and connect the platinum wires with the wires from an electric battery of three cells (Bunsen or Grove). What is the effect?



FIG. 3.

Pour the water out of the bottle and replace it with water to which 1 or 2 cc. of sulphuric acid have been added, and connect with the battery as before. Can you explain the difference?

**\*24.** Collect some of the gas given off in a test tube, by first filling a small test tube completely with water, then inverting it over a dish of water with the open end below the surface of the water, and allowing the gas bubbles to rise under the open end of the tube; when the tube is full of the gas, cover the end with the thumb and bring the end of

the tube to the flame of a burner; remove the thumb. Does the gas burn?

(It is customary to observe whether or not a gas burns, and this experiment teaches that it is well to perform such experiments with care when dealing with unknown gases.)

**\*25.** Collect some of this same gas in a eudiometer tube filled with *dry* mercury inverted over a bath of mercury; pass the gas first through a tube containing calcium chloride. Why?

Do not use more than 10 cc. of the gas in the tube; bring the open end of the tube down upon a large cork under the mercury, to serve as a cushion, and clamp the tube firmly in position; pass an electric spark through the gas by means of a battery and Ruhmkorff's coil. Do not stand too close to the tube when the spark is passed, so as to avoid accident in case the tube breaks.

Note carefully what takes place; examine the tube afterwards. Can you observe any moisture upon it?

**\*26.** Is the gas that escapes from the two platinum plates the same in both cases?

Arrange the platinum plates in an open dish under water, in such a manner that a test tube filled with water may be inverted over each plate, and thus collect the gas that escapes from each plate. (See Fig. 4.) Does the same amount of gas escape from each? To determine this, equalize the pressure in the tubes, and mark accurately on each the point to which it is filled with gas; when it is emptied later, measure out from the burette the exact amount of water required to fill the tube to the marked point, thus learning the exact volume of gas that was present in each tube. Before emptying the tube, examine the gases as to appearance, color, etc. Close the tube containing the most gas with the thumb while its open end is still below the surface of the

water, invert it, and bring a lighted match to the mouth of the tube and remove the thumb; is the result the same as in Exp. 24? What do you observe?

In the same manner examine the gas in the other tube. Note carefully the difference between the two. Explain why the gas collected in Exp. 24 was different from either of these.

The gas that is present in the greatest proportion is called *hydrogen*, and the other gas is known as *oxygen*. We have seen that water contains hydrogen and oxygen. Does it contain anything else? Why?

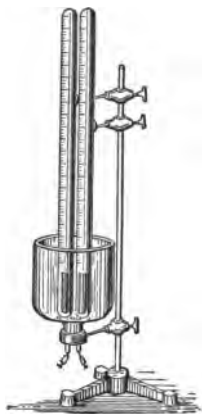


FIG. 4.

### OXYGEN.

Let us study these gases more carefully before going farther with water, first taking up oxygen. Oxygen may be obtained in a number of ways other than by the action of electricity upon water.

**27.** Put some mercuric oxide in a hard glass tube, close the tube with a cork carrying a delivery tube, and heat it very hot, using a blast lamp, if necessary; collect the gas given off in a test tube over water. Examine the gas. What is it? Examine the tube and note the change that has taken place in the mercuric oxide.

**28.** Take a little black oxide of manganese, and heat as indicated in the previous experiment. It will be necessary to heat the tube to redness. Is oxygen obtained?

**29.** Take a little potassium chlorate, heat in a test tube, collect the gas given off; observe carefully the behavior of the potassium chlorate. What is the gas?

**30.** Mix intimately in a piece of paper, with a knife, equal parts of potassium chlorate and manganese dioxide, about 10 grams of each; heat in a test tube. Examine the gas. How does this differ from the previous experiment?

**31.** Mix intimately 30 grams of potassium chlorate and 20 grams of manganese dioxide;<sup>1</sup> put the mixture into a glass retort and heat; pass the gas given off into a gasom-

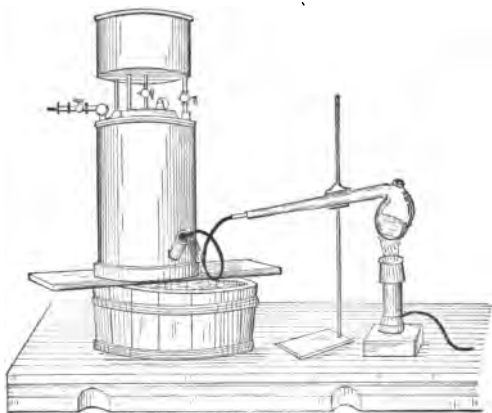


FIG. 5.

eter;<sup>2</sup> heat as long as gas is given off. When no more gas is given off, remove the delivery tube from the gasometer,

<sup>1</sup> It sometimes happens that manganese dioxide is adulterated with powdered coal, or is mixed with dust and quantities of organic matter, so that when mixed with potassium chlorate and heated, there will be an explosion. To avoid danger from this source, a little of the manganese dioxide and potassium chlorate should first be heated in a test tube, as in Exp. 30, to learn whether or not the manganese dioxide is sufficiently pure to use for this purpose.

<sup>2</sup> Where large gasometers are not available, each student may make a very satisfactory one for himself as follows: Take a large flask, 2, 3,

then remove the heat from the retort, screw the cap on the gasometer, and reserve the gas therein for farther experiment. (See Fig. 5.)

Try to light a jet of oxygen as it escapes from the gasometer. Does it burn?

Collect a small cylinder or jar full of oxygen from the gasometer, by first filling the cylinder with water, and then inverting over water, allowing the gas to pass under the open mouth of the jar below the surface of the water as shown in Fig. 7.

Place in the jar a moist piece of red and a moist piece of blue litmus paper. Has the oxygen any effect upon either?

4, or 5 liters, according to size desired, fit to this a two-holed rubber stopper, through one hole pass a glass tube extending to the bottom of the flask and bent at right angles just outside the stopper; to the outer end of the tube attach a rubber tubing 45 cm. long, and close with a pinchcock. Through the other hole pass a tube extending just through the stopper, bent at right angles, also closed with rubber tube and pinchcock. (See Fig. 6.) See that all joints are air-tight. Connect the rubber tubing *A* with the hydrant, open the pinchcocks, and allow water to run in until

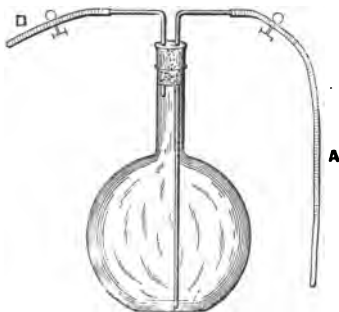


FIG. 6.

full. Close pinchcock *A* and disconnect from the hydrant. Connect *B* with the vessel generating the gas; open pinchcock *A*, and as the gas flows in, water flows out; when full, close both pinchcocks and disconnect the generator. The gas may be obtained from this flask as desired by connecting *A* with the hydrant and allowing water to run into the flask, forcing the gas out at *B*.

Lower into it a burning candle; does oxygen support combustion? Is the litmus paper changed?

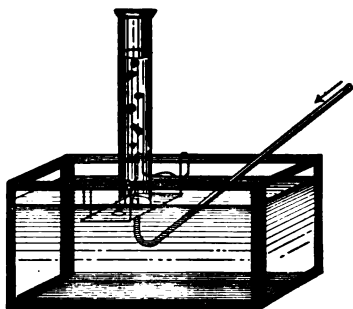


FIG. 7.

(Litmus is a vegetable coloring matter; litmus paper is made by dyeing paper with this coloring matter.)

Some substances have the power of turning red litmus blue; such substances are said to have an *alkaline reaction*. Other substances turn blue litmus red; substances of

this kind are said to have an *acid reaction*. Still other substances do not affect litmus one way or the other; these substances are said to have a *neutral reaction*.

**32.** What effect does oxygen have upon other substances? Take three hard glass tubes about 20 cm. long and 12 to 15 mm. internal diameter; after rounding the sharp edges on the ends of the tubes, fit carefully into each end a cork. Let a straight tube of glass 5 or 6 cm. long pass through one cork; upon this tube is a rubber tube 5 or 6 cm. long closed with a pinchcock; the other cork is perforated by a hole carrying a glass tube about 20 cm. long, bent at right angles near the cork, and extending under the surface of water in a beaker, thus making the tubes air-tight. (See Fig. 8.)

In one of these tubes place a piece of charcoal about the size of a large pea; in another, a piece of sulphur; and in the third, a piece of phosphorus of the same size.

NOTE. In working with phosphorus, use the greatest care, as it is liable to take fire and produce very painful and dangerous burns.



It is always kept under water. When a small piece is desired, transfer a stick of it to an evaporating dish filled with water, and holding it with a pair of forceps, cut off what is wanted *while it is under the water; never touch it with the hands.* Do not leave small pieces of it lying about, but take care that all that you do not use is returned to its proper place in the bottle.

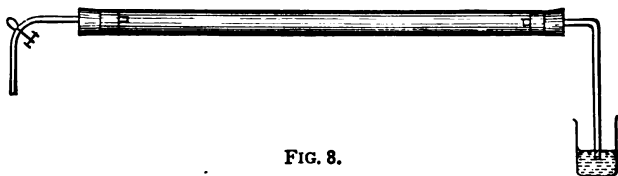


FIG. 8.

Connect each tube in turn with the gasometer, open the pinchcock, and allow oxygen to drive out and replace the air in each tube. Then close the pinchcock and allow the tubes to stand for an hour. At the end of this time examine carefully and note any changes that have occurred.

**33.** After each tube in the previous experiment has stood at least one hour, and has then been examined, take the tube containing the piece of charcoal and connect with the gasometer; allow oxygen to pass slowly through. Heat the tube about the charcoal and note the changes. Finally, treat the tubes containing the sulphur and phosphorus in the same manner and in the order named.

Have at the end of each tube, so that the products formed by the action of oxygen upon the other substance shall pass over them, two pieces of moist litmus paper, one of each color.

In each case examine the water through which the gases have passed; where possible collect the gas coming through in a test tube, and examine it. Is it oxygen? What change, if any, do you observe in the water? What becomes of the oxygen? (See note after Exp. 35.)

**34.** Fill a liter flask with oxygen; invert the flask, and quickly insert a burning candle attached to a bent wire (see Fig. 9), and put the mouth of the flask under water, having previously placed in the water a little of a solution of caustic soda. Explain the phenomena observed.

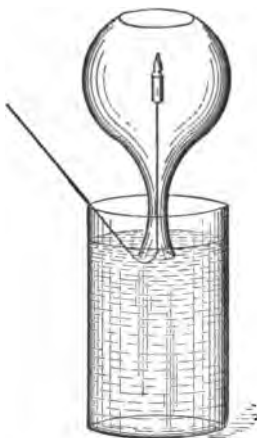


FIG. 9.

**35.** Take a piece of tube about 20 cm. long and 15 mm. internal diameter, similar to the one used in Exp. 32; fit corks and tubes to the ends of it as in that case, except that here both tubes are short and straight. Put a loose plug of asbestos about 5 cm. long next to one of the corks; carefully weigh the tube with the plug and corks; then put into it about 40 cm. of

magnesium wire loosely coiled, and weigh again. Connect a calcium chloride tube with the end of the tube farthest from the asbestos plug, and connect the calcium chloride tube with the gasometer containing oxygen; pass a *very slow* stream of oxygen through the tube, and heat the tube about the magnesium wire until it is nearly red hot. After all the wire has burned, allow the apparatus to cool, and weigh. Examine carefully the substance left in the tube.<sup>1</sup>

**36.** Repeat the last experiment, using air instead of oxy-

---

<sup>1</sup> The heat of the burning magnesium frequently cracks the glass tube; it is therefore well to protect the tube by putting a thin layer of loose asbestos along the bottom of the tube, and allowing the magnesium to rest upon that. The same difficulty occurs with the burning phosphorus in Exp. 33, and may be avoided in a similar manner.

gen. What differences do you observe? What conclusion does this enable you to draw in regard to the air?

**37.** Prepare a tube as in Exp. 35, omitting the asbestos plug, fill it  $\frac{2}{3}$  full of copper turnings, and weigh carefully; fill the tube with dry oxygen and connect with a drying tube, which in turn is connected with a liter flask filled with oxygen. Through one hole of the cork closing the flask a glass-stoppered separating funnel should extend (see Fig. 10); by filling the funnel with water and allowing

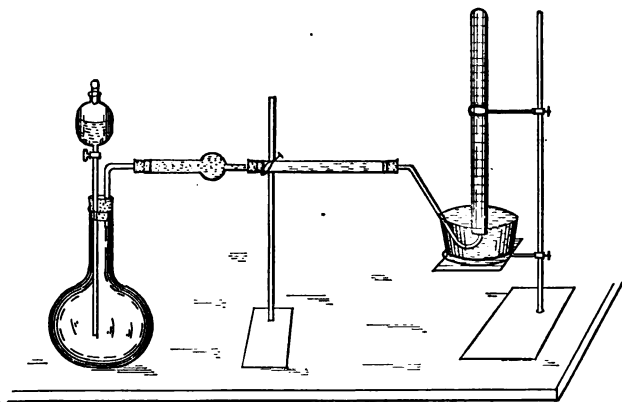


FIG. 10.

it to flow *slowly* into the flask, the oxygen is driven out through the tube containing the copper, which should be heated to redness. Any gas that passes through the apparatus should be collected in a cylinder over water, its volume accurately determined, and then it should be examined to determine whether or not it is oxygen.

After nearly all the oxygen has been driven through the apparatus, the operation is stopped, and the whole allowed to cool. The tube containing the copper is then weighed,

and its change in weight accurately noted. The volume of water poured into the flask should be carefully measured; this is equal to the volume of oxygen driven into the apparatus. This volume less the volume collected in the cylinder at the end of the apparatus is the volume of oxygen which has disappeared; what has become of it? How many cubic centimeters of oxygen have disappeared? What has been the change of weight in the copper? How much does the oxygen which has disappeared weigh? From this calculate what a liter of oxygen would weigh under standard conditions.

**38.** Another method for determining the weight of a liter of oxygen is as follows: Arrange an apparatus like the

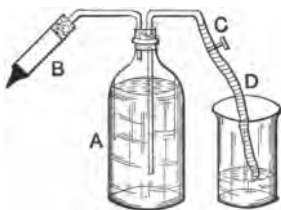


FIG. 11.

one shown in Fig. 11. Fill the flask *A*  $\frac{3}{4}$  full with distilled water; weigh the small hard glass tube *B* accurately and then put into it about 0.3 of a gram of potassium chlorate and weigh again. Put the tube in place and convince yourself that the apparatus is air-tight by opening the clip *C* and

blowing into the flask with all your power; then closing the clip before removing the mouth, allow the flask to stand for some moments; if then, upon opening the clip, water flows out of the tube, you may be sure that the apparatus is air-tight. Allow the delivery tube *D* to be full of water, and assure yourself that the air in the neck of the flask is under the atmospheric pressure. Why? Close the clip; squeeze the water out of the rubber tube below the clip. Place a clean and perfectly dry beaker under the rubber tube, open the clip, and begin to heat the hard glass tube containing the potassium chlorate. Heat as long as any gas is evolved.

When no more gas is given off, remove the flame and allow the apparatus to cool, taking care to have the end of the rubber delivery tube below the surface of the water that has been driven out. Why? After the apparatus has cooled down to the temperature of the room, bring the air and oxygen in the flask to the atmospheric pressure by bringing the water in the beaker to the same level as that in the flask; close the clip; squeeze out the water in the rubber tube and measure accurately the volume of water which has been driven out; this is the volume that the oxygen generated occupies at the temperature of the room and the atmospheric pressure. Find the volume it would occupy under standard conditions.

Remove the tube which contained the potassium chlorate, and weigh it; the loss of weight is the weight of the oxygen which has escaped, the volume of which you know. From these data calculate what one liter of oxygen would weigh.

Oxygen forms what per cent of potassium chlorate?

How many grams of oxygen can be obtained from 10 grams of potassium chlorate?

How many liters of oxygen would that be?

One liter of hydrogen under standard conditions weighs 0.0896 gram; how many times heavier is oxygen than hydrogen?

39. Give a summary of your work upon oxygen and the facts that you have learned about it.

## HYDROGEN.

Besides being formed by the action of electricity upon water, hydrogen may be obtained in other ways.

\*40. Obtain a piece of iron pipe 40 to 50 cm. long and about 1 cm. internal diameter (gas pipe answers the pur-

pose), and fill it  $\frac{2}{3}$  full of coarse iron filings or turnings. Connect one end of the tube with a flask containing water, and the other with a delivery tube of glass or rubber, so that any gas formed may be collected over water. Arrange the apparatus as shown in Fig. 12. Heat the iron tube to redness, boil the water in the flask, and collect several samples of the gas given off. What is it? How do you know? Allow the tube to cool; shake out the iron filings

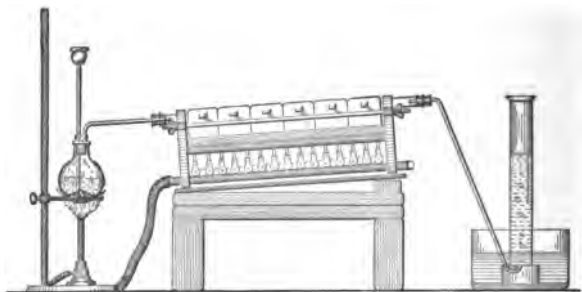


FIG. 12.

and examine them. Explain the changes that have taken place.

**41.** Fill an evaporating dish, 15 to 20 cm. in diameter,  $\frac{2}{3}$  full of water. Take a piece of sodium (sodium is an element) from the bottle with a pair of forceps; free it from its adhering oil by pressing it in filter paper; with a pocket knife cut off three small pieces no more than  $\frac{1}{2}$  the size of a small pea. Then put the main part of the sodium back in the bottle that it came from, taking care to keep it entirely free from water or moisture. Throw one of the small pieces upon the surface of the water.

(*Caution.* As there is frequently a slight explosion here, it is well not to stand too near the dish.)

After this has disappeared, throw a second piece upon the water, and bring a lighted match in contact with it.

Fill a large test tube with the water from the dish, and invert it with the mouth below the surface of the water. Wrap the third piece of sodium in a small piece of filter paper, and, by means of the forceps, bring it quickly below the mouth of the tube, and allow it to rise within the tube. If necessary, introduce a second small piece of sodium in the same way until enough gas has been collected to be examined, in order to find out what gas it is. Is it oxygen or hydrogen? Where did it come from?

Examine the water that is left in the dish, — its feel, its taste, its action on red litmus paper. Evaporate a portion of it, and examine the residue. Is it sodium? Explain what has taken place.

**42.** Take a small piece of potassium (also an element) about  $\frac{1}{2}$  or  $\frac{1}{4}$  the size of a pea, and drop it upon the surface of water. Observe caution as in the previous experiment. What difference do you observe in the action of potassium upon water?

The changes are similar to those in the case of sodium.

**43.** The most convenient method of preparing hydrogen in the laboratory is by means of zinc and hydrochloric acid.

Place about 20 grams of granulated zinc in a flask of about 500 cc. capacity; fit to this flask a cork through which are bored two holes. Through one of the holes insert a funnel tube; in the other fix a glass delivery tube. (See Fig. 13.) Pour into the flask, through the funnel tube, about 100 cc. of dilute hydrochloric acid. What is the effect?

*After a few moments* collect some of the gas given off in a test tube over water. What effect has the gas upon litmus paper? Fill, over water, five jars or bottles with the gas, and set them aside for further examination. Invert the test

tube filled with gas, and bring a lighted match quickly to its mouth. Is the gas hydrogen?

Replace the delivery tube in the cork by a short jet tube. Wrap a towel carefully around the flask. Why? (See note after Exp. 47.) Bring a lighted match to the jet. Has the gas any odor? Color? Taste? Where does the gas come from? (Zinc is an element.)

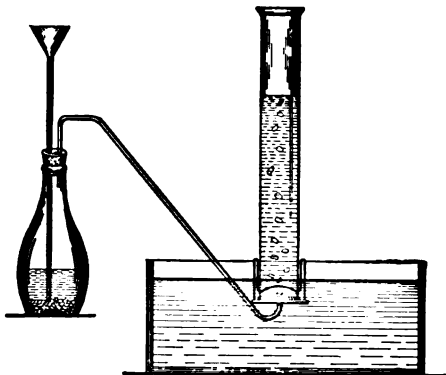


FIG. 13.

**44. Does hydrogen support combustion?**

Light a candle, raise one of the jars of hydrogen slightly out of the water, and with the mouth of the jar still down, introduce the candle.

**45. Which is the heavier, air or hydrogen?**

Take two of the jars that are filled with hydrogen; place one with its mouth upward and open, the other with its mouth downward and open, and after a moment bring a lighted candle into each. What is the difference? Explain the difference.

**46.** Take one jar from which the hydrogen has escaped, and another which is filled with hydrogen; pour the hydro-



gen from one vessel into the other. How must the receiving vessel be arranged? How do you prove that the hydrogen gas has gone from one vessel to the other? What does this experiment show?

47. Arrange a flask containing zinc as in Exp. 43. Instead of a delivery tube in the cork, put a short tube bent at right angles. Connect this with a tube filled with calcium chloride, and connect the end of the calcium chloride tube with a glass jet; then pour dilute hydrochloric acid through the

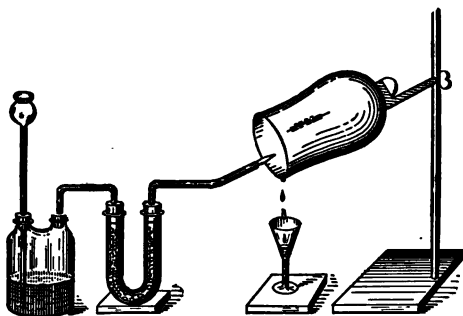


FIG. 14.

funnel tube after the action has been going on for seven or eight minutes; wrap the flask with a towel, and light the hydrogen escaping from the jet. (See note following this experiment.) What is the effect of the calcium chloride upon the hydrogen? Place a large bell jar over the flame. (See Fig 14.) After a few moments examine the surface of the bell jar. What do you find? Explain this. What chemical change takes place when hydrogen burns?

NOTE. Owing to the danger of lighting the hydrogen before all the air is removed from the flask, it is well to test a small sample of the gas before lighting it, to learn whether or not all the air is removed. This can be done by collecting a test tube full of the gas over water, cover-

ing the tube with the thumb, and bringing it to the flame with the mouth downward; remove the thumb. If the hydrogen ignites with a slight report, and then continues to burn quietly upon inverting the tube, it will be safe to light the gas jet; if, on the other hand, the gas ignites with a sharp report, the apparatus still contains air, and the gas should be allowed to run some time, and be tested again before lighting.

*Never, under any circumstances, omit to wrap the generator in a towel, so that if it does explode, no serious accident will result.*

48. Take a small flask and put in it a little iron (iron filings or scraps or nails, etc.); add to this a little hydro-

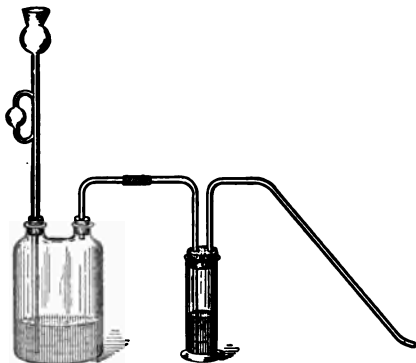


FIG. 15.

chloric acid. Examine the gas given off; is it hydrogen? Has it odor? Pass the gas through a wash bottle containing a solution of potassium permanganate to which a little sulphuric acid has been added. (See Fig. 15.) What effect does this have upon the odor? Can you explain the cause of this?

A number of other metals give hydrogen when hydrochloric acid is added to them; for example, tin, magnesium,

and aluminium. Where does the hydrogen come from? These metals are all *elements*.

49. If you had used sulphuric acid instead of hydrochloric acid, would you have obtained hydrogen? Try it.

Acetic acid gives the same results as sulphuric acid. What conclusion does this enable you to draw in regard to the composition of hydrochloric acid, sulphuric acid, and acetic acid?

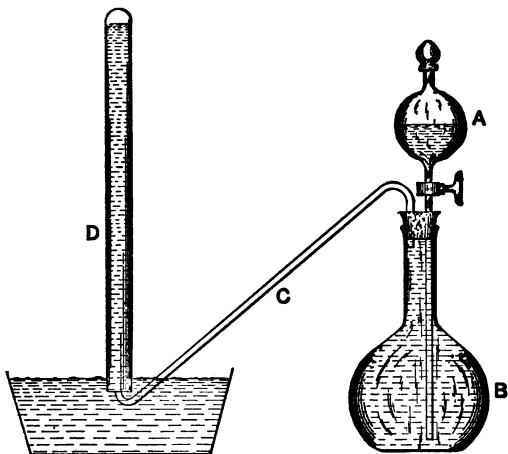


FIG. 16.

50. Does a given amount of zinc always give, with sulphuric acid, the same amount of hydrogen? If so, how much?

Weigh out accurately a small piece of zinc not more than 0.2 of a gram; take the apparatus prepared for the purpose. (See Fig. 16.) Put the zinc in the flask and fill with water by pouring the water into the large funnel tube *A* until the flask *B* and delivery tube *C* are completely filled with

water. Allow the delivery tube to extend into a basin of water and bring over the mouth of it a graduated tube *D* capable of holding 100 cc. (a 100 cc. burette will answer), filled with water. Now pour dilute sulphuric acid into the tube *A*, and allow it to run into the flask and come into contact with the zinc. The hydrogen will pass out of the delivery tube and be collected in the graduated tube. After all the zinc is dissolved, run water through the apparatus until all the hydrogen is transferred to the graduated tube; then carefully transfer the graduated tube to a cylinder of water and measure the volume of the gas, bringing the level of the water inside the tube to the level of the water in the cylinder. Why? Note the temperature and the height of the barometer, and reduce the volume to standard conditions.

Calculate from these data the volume of hydrogen one gram of zinc would give.

**51.** Repeat Exp. 50, using hydrochloric acid instead of sulphuric acid; in this case it is well to have a little caustic potash in the cylinder to which the graduated tube is transferred. What volume of hydrogen is obtained by using one gram of zinc? Is it the same as from sulphuric acid?

Very careful experiment has shown that one liter of hydrogen under standard conditions weighs 0.08958 gram. What *weight* of hydrogen will one gram of zinc produce when dissolved in hydrochloric acid? What weight of zinc is required to produce one gram of hydrogen?

It always requires the same amount of zinc to produce one gram of hydrogen. This number is the *equivalent weight of zinc*. The number representing the equivalent weight of zinc is the number of parts of zinc that are required to produce one part of hydrogen. The same is true of the equivalent weight of the other metals.

**52.** Repeat Exp. 50, using magnesium instead of zinc. Weigh out about 0.1 gram of magnesium.

What is the equivalent weight of magnesium? What weight of hydrogen could be obtained by dissolving one kilogram of magnesium in sulphuric acid? In hydrochloric acid? How many liters of hydrogen in each case? What weight of magnesium would be required to produce one gram of hydrogen?

**53.** Find the equivalent weight of tin by using tin and hydrochloric acid as in the previous experiment. Answer the same questions in regard to tin that were answered for magnesium.

**54.** Take a piece of glass tube about 25 cm. long and at least one cm. in internal diameter, and fit a cork to one end. Pour about 10 cc. of water into an evaporating dish, and add enough plaster of Paris to form a thick cream, stirring it thoroughly with a glass rod. Before it sets withdraw the cork from the tube, and plunge the end of the tube into the plaster of Paris, until it rests upon the bottom of the dish, and support it in its position until the plaster is moderately hard. Then carefully detach it from the surrounding plaster, and push the cork a little way into the tube, so as to drive the plug of plaster before it. Take the cork out, and dry the plaster thoroughly by a gentle heat, leaving it in a warm place for several hours. When it is quite dry, insert the cork and fill the tube with dry hydrogen by displacing the air. When the tube is full, bring the lower open end into a vessel of water, and clamp it in position; remove the cork from the upper end. Observe carefully what change takes place. Does a tube filled with air instead of hydrogen behave in the same way?

**55.** Allow the tube used in Exp. 54 filled with air to stand with the cork removed and the lower open end under water.

Carefully bring over the tube a large bell jar filled with hydrogen. What is the effect here? Can you offer any explanation of the phenomena observed?

56. Make a summary of your work upon hydrogen.

57. We have seen that two volumes of hydrogen combine with one of oxygen to form water. Knowing how much heavier oxygen is than hydrogen, it is a simple matter to calculate the proportion *by weight* in which oxygen and hydrogen combine in water. What is it?

This can also be determined by direct experiment in a number of ways.

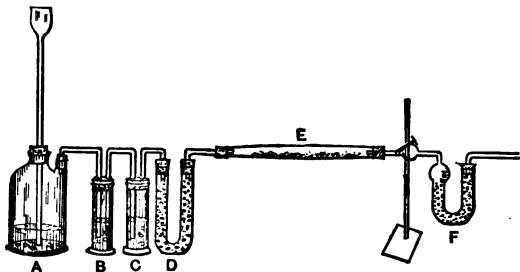


FIG. 17.

58. Arrange an apparatus like the one shown in Fig. 17. The flask *A* contains zinc; the bottle *B*, a solution of potassium permanganate; the bottle *C*, concentrated sulphuric acid; and the tube *D*, granulated calcium chloride. What is the object of each of these? In the tube *E* place some copper oxide; the tube *F* contains granulated calcium chloride.

Disconnect the flask *A* from the rest of the apparatus; connect the filter pump with tube *F*, and slowly draw dry air through the apparatus, heating the tube *E* nearly to redness to drive out all moisture; allow it to cool with the current

of dry air still passing through it. When completely cool, disconnect the tube, and weigh it accurately; weigh also the tube *F*. Connect the apparatus again, as shown in the figure; pour dilute sulphuric acid through the funnel tube into the flask *A*; after the hydrogen has been passing through the apparatus five or six minutes, collect a sample that issues through the tube *F* in a test tube, and examine it. When convinced that all the air has been expelled from the apparatus, heat the copper oxide in the tube to redness; carry on the operation for fifteen minutes, keeping a gentle current of hydrogen passing through the apparatus, and the copper oxide red hot. Note the changes. Remove the heat, allowing the hydrogen still to pass through the apparatus until cool; then disconnect the flask *A*, and draw air through the apparatus as before to displace the hydrogen. Then weigh tubes *E* and *F* again, noting their changes of weight. Explain to what this change is due. Explain how these weights give us any clue to the composition of water by weight. Make the calculations from your data, showing the percentage composition of water, also the amount of oxygen as compared with the amount of hydrogen as unity. How does this result agree with the result calculated from your other data? Explain the differences. What weight of oxygen and what weight of hydrogen can be obtained by electrolysis from one kilogram of water? What *volume* of each gas can be obtained in this way from one kilogram of water? Calculate the weight of a liter of oxygen from this data, and compare the result with the result obtained by actual experiment.

**\*59.** Determine the weight of a liter of steam. The apparatus for this purpose consists of a wide tube about 15 mm. internal diameter, sealed at one end, and graduated in cubic centimeters, and capable of holding about 100 cc.

This tube is filled with mercury, and inverted over a dish of mercury; the upper portion of this tube is surrounded with a larger tube of glass containing a thermometer, into which the vapor from the boiling liquid in flask *A* is passed. (See Fig. 18.) Procure a little bottle made especially for

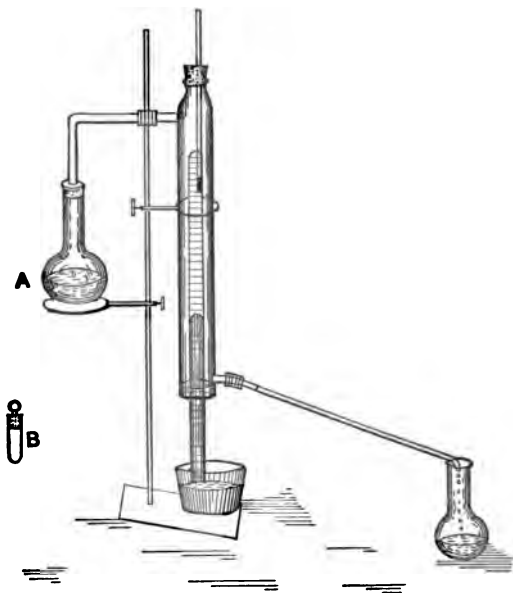


FIG. 18.

the purpose (*B* in the figure), have it perfectly clean and dry, and weigh it accurately; fill it with water, and weigh again. (The weight of water should not exceed 0.07 gram; if the bottle holds more than that, do not fill it.) Bring the weighed bottle of water below the open end of the tube under the mercury, and allow it to rise in the tube to the top of the mercury. Boil the liquid in flask *A*. (This



should be a liquid the boiling point of which is above  $100^{\circ}$ ; aniline boiling at  $185^{\circ}$  answers very well; mesitylene ( $163^{\circ}$ ), nitro-benzene ( $205^{\circ}$ ), xylene ( $140^{\circ}$ ) will also answer. The vapor is condensed and collected, as shown in figure, to avoid loss.) The temperature in the jacket gradually rises to the boiling point of the liquid used; the water is converted into steam, and its volume may be read off. The pressure and temperature at which the volume is measured should also be observed, in order that it may be reduced to standard conditions. Calculate the weight of a liter of steam.

As a matter of fact, could steam be measured under standard conditions?

What is the specific gravity of steam referred to hydrogen? Hydrogen is frequently used as the unit of specific gravity for gases, just as water is used as a unit for liquids.

How many liters of steam would be formed from one kilogram of water?

How does this volume compare with the volume of hydrogen formed from a kilogram of water? With the volume of oxygen formed from a kilogram of water?

When one volume of oxygen unites with two volumes of hydrogen to form water, how many volumes of steam does this water form?

The results calculated in this way are found by experiment to be quite correct.

60. Summarize your work upon water.

### HYDROCHLORIC ACID.

61. Arrange an apparatus like the one shown in Fig. 19. In the flask *A* put 60 grams of coarse, common salt; pour upon this, through the funnel tube *B*, some dilute sulphuric acid, prepared by pouring 50 cc. of concentrated commer-

cial sulphuric acid into 30 cc. of distilled water, and allowing the mixture to cool. (Owing to the great heat caused by mixing sulphuric acid with water, the mixture should be made in a beaker, or other vessel not broken by heat; *the*

*acid should be poured into the water*, for when the water is poured into the acid, some of it is liable to be suddenly converted into steam, throwing the acid out of the dish.)

Heat the flask containing the salt and sulphuric acid gently; at first the air will be driven out of the flask, then a gas escapes that very evidently is not air. This gas is called *hydrochloric acid gas*. Fill three jars (or wide-mouthed bottles) with the gas, by allowing the delivery tube to extend to the bottom of the jar, which is loosely covered with a glass plate; when filled withdraw the

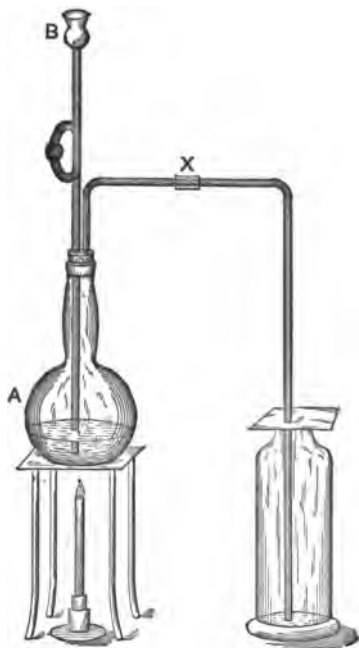


FIG. 19.

delivery tube, and cover tightly with the glass plate, and set aside. After two bottles are thus filled, try to light the gas as it escapes from the delivery tube. Does it burn? Then connect the flask at X with three Wolff bottles, as shown in Fig. 20, the first bottle carrying a thermometer, and all three about  $\frac{1}{8}$  full of distilled water. While the gas is running into

the water, examine the bottles of gas already collected. Note the odor and color of the gas. Remove the glass plate from one of the bottles, and blow your breath over the mouth of the bottle. What effect has this? Drop a moist piece of blue litmus paper into the gas. What effect has the gas upon it? What, therefore, is the nature of the gas?

Lower a burning candle or splinter into the second jar of gas. Does it support combustion? Take the third jar of gas, holding the glass cover in position, invert the jar, and bring

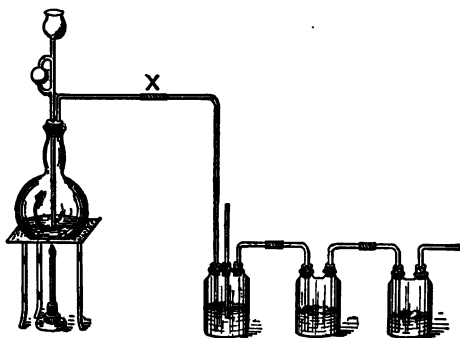


FIG. 20.

its mouth under the surface of some distilled water in a beaker; remove the glass plate. What occurs? Explain it.

Examine the water by testing it with blue litmus paper; taste of it. Why cannot this gas be collected over water as hydrogen and oxygen were? Is hydrochloric acid gas heavier or lighter than air? Your method of collecting it will answer this question.

**62.** After a constant stream of gas has passed into the water in the Wolff bottles for  $\frac{1}{2}$  or  $\frac{3}{4}$  of an hour (it may be necessary to renew the charge of salt and sulphuric acid in the generating flask), remove the heat, and disconnect the flask

at *X*. Save the contents of the generating flask for examination as directed in Exp. 64. What has been the effect of the gas upon the temperature of the water? Examine some of the water in the first Wolff bottle. What is its effect upon litmus paper? Drop a piece of metallic zinc into a test tube containing some of this solution. What happens? What gas is given off?

Compare the water solution of hydrochloric acid you have made with some of the laboratory hydrochloric acid. Are they the same? Find the specific gravity of the solution in each of the three Wolff bottles, using for this purpose a hydrometer. Which is the heavier? What would be the weight of a liter of the liquid in each bottle? Explain how the hydrometer gives the specific gravity.

63. Fill a burette with a solution of sodium hydroxide. What effect has sodium hydroxide upon litmus paper? Allow the sodium hydroxide in the burette to run slowly into a beaker containing 50 cc. of the strongest hydrochloric acid you have prepared. Stir the acid with a glass rod while the hydroxide is running in; allow the liquid from the burette to run into the acid until the hydrochloric acid has just lost its power to redden blue litmus paper, being careful not to add more of the sodium hydroxide than is required for this purpose.

Evaporate this mixture almost to dryness in a small evaporating dish, and then allow to stand until crystals are formed. Take some of these crystals, dry them, examine them carefully, taste them, put some strong sulphuric acid upon some of them in a test tube. What gas is formed? What are the crystals?

How can sodium hydroxide be made? (Ref. to Exp. 41.) Sodium hydroxide is a type of a certain class of bodies that have the power of *neutralizing* or destroying acid properties.

Such bodies are called *bases*. Conversely, when we add an acid to a substance like sodium hydroxide, the *basic* power of the hydroxide is neutralized, or destroyed. The properties of the two become merged in a new substance, and can no longer be recognized. The substance, which we have just seen is produced by neutralizing the *base*, sodium hydroxide, with hydrochloric acid, is a type of a large class of such neutralization products. These products are called *salts*.

†64. After the substance in the generating flask used in Exp. 61 has cooled down, add about 200 cc. distilled water, and after all has gone into solution, pour out into an evaporating dish, evaporate the substance to dryness, and heat until all white fumes of sulphuric acid are given off; then cool, and add just enough distilled water to redissolve the substance; when in solution, evaporate off about  $\frac{2}{3}$  of the water, and allow to stand until crystals form. Pick out a few of these crystals and examine them: are they crystals of common salt? When heated with a little strong sulphuric acid in a test tube, do they give off hydrochloric acid? Allow a few of the crystals to lie on a piece of filter paper exposed to the air. What are the crystals?

\*65. In Exps. 43 and 51 we have seen that hydrochloric acid is not a chemical element, for it contains hydrogen. The volume of hydrogen contained in a definite volume of hydrochloric acid gas may be determined as follows: Into the open limb of the U tube (Fig. 21) mercury is poured until both limbs are full, the air being driven out of the right-hand limb through the open stopcock: the stopcock is then connected with the generating flask, delivering pure,

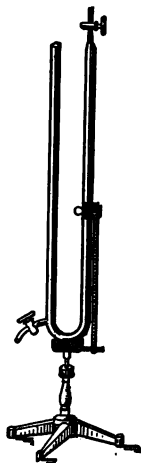


FIG. 21.

dry hydrochloric acid gas. The pinchcock at the bottom of the U is then opened. As the mercury flows out, the pure gas enters and fills the right-hand tube. Adjust the level of the mercury to the same height in both tubes, and measure the volume of gas by putting a label on the tube so that its upper edge marks the position of the mercury meniscus, in case the tube is not graduated. Mercury is added to the open arm, the stopcock being previously tightly closed, until it is within 5 cm. of the top; the remainder of the tube is completely filled with sodium amalgam, made by pressing several small pieces of clean, freshly cut sodium one by one under the surface of 5 or 6 cc. of mercury in a porcelain mortar.

The open end of the tube is closed with a rubber stopper, and is inclined and shaken so as to bring the hydrochloric acid gas in contact with the sodium amalgam; after a few moments' shaking, the gas is transferred to the closed limb of the tube, the rubber stopper removed, the mercury brought to the same level in both limbs, and the new volume measured. That the gas remaining is hydrogen can be shown by again filling the open arm with mercury, opening the stopcock, collecting the escaping gas over water, and examining it.

**\*66.** The other constituent of hydrochloric acid may be obtained by decomposing hydrochloric acid with electricity. The apparatus for this purpose consists of two stopcock burettes, *A* and *B*, provided with carbon electrodes *a'*, *b'*: the burettes are connected with the reservoir tube *C*. (See Fig. 22.) Fill the apparatus with a saturated solution of common salt, containing  $\frac{1}{10}$  its volume of strong hydrochloric acid solution, connect *a'* and *b'* with a battery of 6 to 8 Bunsen elements. After a few moments examine the gases. Which one is hydrogen? Observe the color and

odor of the other. This gas is an element; it is called *chlorine*. After the battery has been running some little time, allow all the gases to escape, close the stopcocks and allow to run for 10 or 15 minutes longer, and observe the relative volumes of the two gases. As chlorine is very soluble in water, the volume relations at first obtained are liable to be wrong. Chlorine is much less soluble in strong salt solution than in pure water, and after the current has run some time, the liquid becomes saturated with chlorine, and no more will dissolve; then accurate volume measurements may be made. What are the relative volumes of hydrogen and chlorine in hydrochloric acid?

One liter of hydrogen would unite with how much chlorine? How much hydrochloric acid gas would this produce?

**\*67.** The weight of a liter of hydrochloric acid gas may be found as follows: Take the U tube used in Exp. 65. (See Fig. 21.) Fill the closed arm with dry hydrochloric acid gas over clean dry mercury, equalize the pressure in the two arms and measure the volume, noting the temperature and height of the barometer, so that the volume may be reduced to standard conditions. Fill a Liebig's potash bulb with a solution of sodium hydroxide, carefully weigh it, then connect it with the stopcock by means of a stout piece of rubber tubing, and slowly force the hydrochloric acid into the potash bulb by filling the open arm of the U tube with mercury and opening the stopcock so that the gas slowly escapes into the bulb. After all the gas has

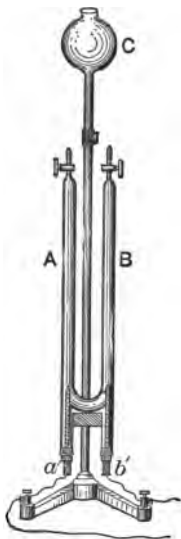


FIG. 22.

been forced into the bulb; it is again weighed. The increase in weight is the weight of the measured volume of hydrochloric acid. Calculate the weight of a liter of hydrochloric acid gas.

How many times heavier than hydrogen is it?

Knowing the weight of a liter of hydrochloric acid and the relative volume relations of the gases composing it, what would be the weight of a liter of chlorine?

### CHLORINE.

†68. Chlorine may be made as follows: Arrange an apparatus like the one used for the preparation of hydrochloric acid. (See Fig. 19.) Into the generating flask put 30 grams manganese dioxide ( $\text{MnO}_2$ ), and pour through the funnel tube hydrochloric acid. (If the flask is charged with equal weights of manganese dioxide and common salt, sulphuric acid may be used instead of hydrochloric acid.) Apply a gentle heat and collect six cylinders of chlorine gas by downward displacement of the air. When each cylinder is full cover it tightly with a glass plate and set aside for further examination. Try to light the jet of escaping chlorine; does it burn? Observe its color and odor. After the cylinders are full, allow the escaping chlorine to run into water while the gas collected in the cylinders is being examined. *Be careful not to inhale this gas.*

†69. Holding the glass plate in position with the hand, invert one of the cylinders of chlorine, bringing its mouth below the surface of water. Is the gas soluble? Is it more or less soluble than hydrochloric acid?

†70. Drop into the second jar some moist pieces of red and blue litmus paper; also some moist flowers of different colors; after allowing to stand a short time, observe the



effect upon the litmus and flowers. What is its effect upon writing done with ordinary ink? First moisten the paper.

†71. Lower into a third jar a burning candle; does it support combustion? Put a small piece of phosphorus (observe precautions, see note in Exp. 32) on a combustion spoon into a fourth jar of chlorine.

†72. Warm carefully a little fresh turpentine in an evaporating dish, moisten a piece of filter paper with it, then suspend it on a wire or with forceps in the chlorine in one of the remaining bottles. What takes place? Turpentine is composed of hydrogen and carbon (an element). What is formed? What becomes of the carbon?

†73. Arrange an apparatus as shown in Fig. 23; the flask *H* is of about 250 cc. capacity. *A* is a chloride of calcium tube. What is it for? The delivery tube is drawn to a point and bent as shown. In the flask are placed zinc and dilute sulphuric acid. What is produced? Allow the gas to escape for some time; finally wrap a towel around the flask

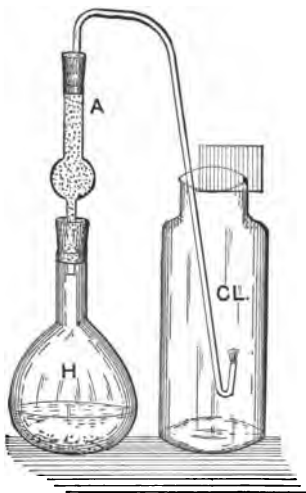


FIG. 23.

and calcium chloride tube; ignite the jet of escaping hydrogen; bring the jet of burning hydrogen into the remaining bottle of chlorine; does it continue to burn? What differences do you observe? What is formed?

†74. Take the generating flask used in Exp. 68. If chlorine is no longer coming off, add a little more manganese

dioxide and hydrochloric acid ; replace the delivery tube by a chloride of calcium tube carrying a short straight tube drawn to a jet. (See Fig. 24.) Collect a jar of pure hydrogen over water. With the mouth of the jar down, light the hydrogen with a match, and bring the gas quickly over the jet of escaping chlorine.

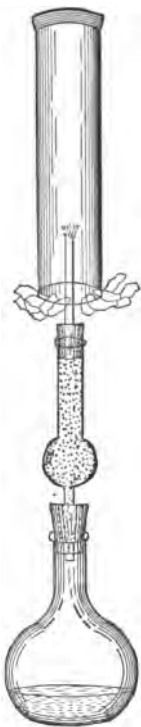


FIG. 24.

Does chlorine burn in an atmosphere of hydrogen? What is formed? Why do you invert the cylinder of hydrogen in this experiment, while in the previous experiment the cylinder of chlorine was not inverted?

**75.** Examine the water into which the chlorine has been passing. What is its odor and color? What is its effect upon litmus paper and upon flowers? Upon writing in ink? Does it affect printing ink?

**\*76.** Arrange an apparatus as shown in Fig. 25. *A* is a hard glass tube upon which a bulb has been blown ; in this bulb are placed some pieces of sodium. *B* is a flask with two tubulations, the tube *K* passing to the hood, or into water containing a little sodium hydroxide. Heat the bulb containing the sodium very gently, and pass a stream of dry chlorine over the molten sodium in the direction of the arrows. Note carefully all the results. When action ceases, allow *A* to cool, and examine the contents of *A* and *B*. What is this substance that has been formed? Knowing that sodium and chlorine are both elements, what is the composition of this new substance that has been formed?

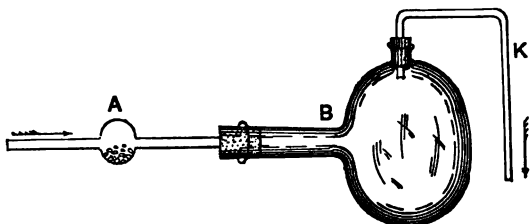


FIG. 25.

**76a.** Summarize your work on hydrochloric acid and chlorine.

### **AMMONIA.**

Ammonia gas is best prepared by heating ammonium chloride (sal ammoniac) with slaked lime.

**77.** Take a little ammonium chloride in a test tube, add a little strong sulphuric acid, and warm. What gas is given off?

**78.** Arrange an apparatus as shown in Fig. 26. *A* is a flask of about 500 cc. capacity. *B* is a cylinder full of quicklime to dry the gas (stick caustic soda may be used; calcium chloride cannot be used because it absorbs ammonia gas). *C* is a cylinder for collecting the gas. Mix thoroughly 25 grams ammonium chloride with an equal weight of air-slaked lime. Place this in the flask *A*, and add just enough water to make pasty, or nearly enough to cover the mixture. Heat the flask gently, and collect four cylinders of gas by displacement of air; cover tightly with a glass plate, and set aside with the mouth down for further examination. Then pass the current of gas into a beaker containing about 100 cc. of distilled water. While this is going on, examine the cylinder of collected gas. Observe color and odor.

Into one of the jars put moist pieces of red and blue litmus paper. To what class of bodies does ammonia belong? Into the same cylinder lower a burning candle or splinter. Does it support combustion? Will it burn? Try to light the dry gas as it escapes from the generator. Invert one of the cylinders full of the gas so that the mouth is up, and try to light it then. Does it burn?

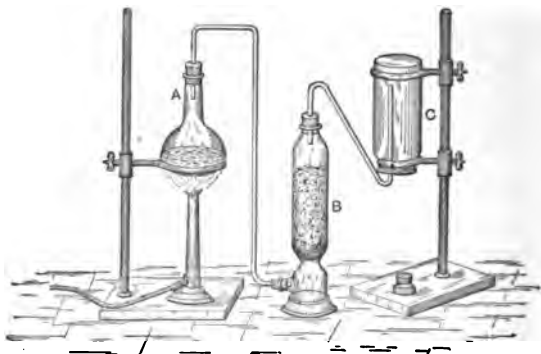


FIG. 26.

Bring the mouth of a second bottle under water, and remove the glass plate. Does it dissolve in water? Examine the water with red and blue litmus paper.

**79.** Fill a cylinder with dry hydrochloric acid gas; cover with a glass plate; bring over this a cylinder of dry ammonia; bring the mouths of the two cylinders near together; remove the plates covering the cylinders, and fit the mouths of the cylinders tightly together. If the edges are ground, so that an air-tight joint can be made, so much the better; then invert the two together so that the cylinder containing the hydrochloric acid is on top.

Observe carefully what takes place. What is formed?

If the joint between the two cylinders is air-tight, they will stick together. Why?

**80.** Examine the water into which the gas has been passing, its odor, its effect upon litmus paper. Find the specific gravity of this solution. Compare this with the specific gravity of the hydrochloric acid solution. Boil some of this solution. What gas is given off?

Take 50 cc. of the laboratory solution of ammonia, and add to this dilute hydrochloric acid until the odor of ammonia just disappears. Evaporate the solution to dryness in a porcelain evaporating dish and examine the residue. Heat a little of it in a test tube with sulphuric acid. What is formed? Heat another portion in a test tube with a little lime. What is formed? Heat a little of it in the lid of a porcelain crucible. How does it behave? What is it?

**\*81.** Electricity decomposes ammonia into two other gases, thus showing that it is not a chemical element. For this purpose, the same apparatus may be used that was used for the decomposition of water (Exp. 26). In this case the apparatus is filled with a strong solution of ammonia, to which a few drops of sulphuric acid have been added, to increase its electric conductivity. Allow the battery to run until sufficient gas has collected in each tube, carefully observing their relative volumes, and the pole at which they are formed. Examine the gas present in larger quantity; its color; odor. Does it burn? What gas is it? Examine the second gas. Does it burn? What is its color and its odor? It is an element called nitrogen. How many volumes of hydrogen unite with one of nitrogen to form ammonia?

**\*82.** Another method of showing the composition of ammonia is, to decompose ammonia with a known volume of chlorine, and measure the nitrogen left over. In this experi-

ment a large tube, prepared for the purpose (see Fig. 27), is filled with a saturated solution of salt, and inverted over a salt solution, and then completely filled with chlorine and allowed to stand for some time until all the chlorine water runs off the sides. The stopcock is then closed, the tube inverted and a little strong ammonia solution is placed in the upper portion of the tube and allowed to enter the chlo-

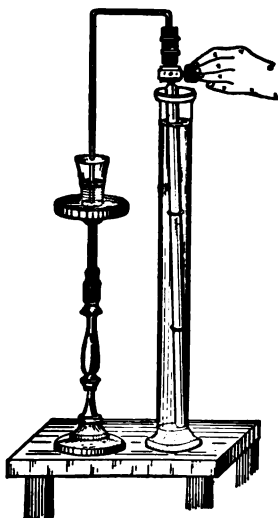


FIG. 27.

rine drop by drop. The chlorine immediately combines with the hydrogen of the ammonia to form hydrochloric acid, setting nitrogen free. The hydrochloric acid combines with some of the ammonia to form the solid ammonium chloride. When all the chlorine has been used up, a little dilute sulphuric acid is allowed to run into the tube to unite with the excess of ammonia that may have been introduced, to form ammonium sulphate. The only gas remaining in the tube is nitrogen. Fill the upper portion of the tube with distilled water; fit a bent tube into a rubber stopper, fill this tube with distilled water; fit the stopper into the open end of the large tube and bring the open end of the bent tube below the surface of water in a beaker. After immersing the long tube in water of ordinary temperature, open the stopcock. If the operation has been carried out without error, the water will flow into the tube until it is two-thirds full, and will then stop. The remaining space is filled with nitrogen under

atmospheric pressure. The arrangement of the apparatus in the last stage of the experiment is shown in Fig. 27. Knowing the volume of chlorine started with, and the volume of nitrogen remaining, how does the experiment give any clue to the volume relations of hydrogen and nitrogen in ammonia? What are the relations, as shown by this experiment?

**\*83.** Take a U tube, like the one shown in Fig. 21, into which are fused platinum wires, so that an electric spark can be passed through it. Introduce over mercury about 25 cc. of pure, dry ammonia gas. (Easily prepared by warming a strong ammonia solution, and drying with lime.) Equalize the pressure in the two arms of the tube and measure the exact volume. By means of a good battery and a Ruhmkorff's coil, pass electric sparks through the gas until it no longer increases in volume. Disconnect the battery; equalize the pressure in the two arms of the tube, and measure the new volume. The electric spark decomposes the ammonia into its constituents, hydrogen and nitrogen.

How many volumes of the mixed gases are obtained from one volume of ammonia gas? How many volumes of hydrogen from one volume of ammonia? How many volumes of nitrogen from one volume of ammonia?

One volume of nitrogen unites with how many volumes of hydrogen to form ammonia? How many volumes of ammonia would be formed?

**\*84.** The weight of a liter of ammonia gas may be determined precisely as the weight of a liter of hydrochloric acid gas was determined in Exp. 67. The U tube is filled with pure, dry ammonia gas, and its volume accurately determined; the ammonia is then forced slowly into a Liebig's potash bulb containing dilute sulphuric acid of known weight. The acid is then reweighed, the increase being the weight of the

known volume of ammonia. Calculate the weight of a liter of the gas.

What is its specific gravity referred to hydrogen?

Knowing the volume relations between ammonia and its constituents, hydrogen and nitrogen, and knowing how many times heavier than hydrogen ammonia is, calculate the specific gravity of nitrogen referred to hydrogen.

How much does one liter of nitrogen weigh?

### NITROGEN.

**85.** In a flask of about 250 cc. capacity put about 10 grams potassium nitrite, 10 grams ammonium chloride, and 10 grams potassium bichromate; over these pour about 30 cc. of distilled water and heat the mixture gently. After the air has been expelled from the flask, collect two cylinders of the gas over water. This gas is nitrogen; does it burn? Is it soluble in water? Note its odor and color.

**86.** Lower into one of the cylinders of nitrogen gas a burning splinter or candle; does it support combustion? Into the second cylinder lower a small piece of burning phosphorus in an iron combustion spoon. Does the phosphorus combine with the nitrogen?



FIG. 28.

**87.** On the surface of some water in a deep basin float, by means of a flat cork, a porcelain crucible cover containing a small piece of phosphorus.

Ignite the phosphorus, and invert over it a large beaker or a bell jar, as shown in Fig. 28.



After the phosphorus ceases to burn, allow to stand until the white fumes disappear. What are these fumes? What becomes of them? Test the water before and after the experiment with blue and red litmus paper. Explain the phenomena you have observed. Test the gas that remains. What is it? Where did it come from?

88. In a hard glass tube, 40 to 50 cm. long, place a roll of bright copper gauze, or fill the tube with copper turnings. Pass a slow stream of *dry* air through the tube and gradually bring it to a bright red heat. Allow the apparatus to run a short time and then collect some of the gas evolved. Examine it and prove what gas it is. Where did it come from?

### THE ATMOSPHERE.

89. Explain the action and use of the barometer.

If a mercury barometer stands at 760 mm., how high will a water barometer stand? A sulphuric acid barometer? An alcohol barometer?

The specific gravity of mercury is 13.6.

The specific gravity of sulphuric acid is 1.83.

The specific gravity of alcohol is 0.82.

What will be the weight of air pressing on a square centimeter of surface when the barometer stands at 755 mm.? At 760 mm.?

90. The air, as we have already seen, contains oxygen, nitrogen, and water vapor; we shall later see that a few other gases are present in very small quantities. Much more than 99 per cent of the atmosphere is made up of the two gases nitrogen and oxygen. The relative proportions in which these two gases are present may be learned as follows:—

Seal one end of a glass tube and fill it  $\frac{1}{8}$  full of cold

distilled water. Invert it in a cylinder of water. Bring the liquid in both tube and cylinder to the same level. Why? Note the volume by making a mark on the tube at the surface of the water. Now take a small piece of

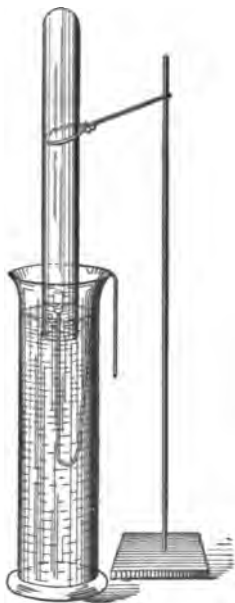


FIG. 29.

phosphorus (see note in Exp. 32), attach a wire to it, and push it up into the tube until it projects into the air space (Fig. 29). Let it stand until there is no further change in volume. Then remove the phosphorus and bring the liquids to the same level, as before, making a mark on the tube at this second level. Cover the open end of the tube with the thumb and invert it. Plunge a lighted stick into it and note what occurs. What does this gas behave like? What has happened to the original air? What has become of the substance removed?

Fill the tube, first to one mark and then to the other, with distilled water from a burette, and note the volume in cubic centimeters in each case. From these measurements calculate the percentage composition of air by volume, assuming it to be composed entirely of oxygen and nitrogen.

Knowing the relative weights of oxygen and nitrogen, calculate the percentage composition of air by weight.

Air is a mixture of the two gases oxygen and nitrogen in the proportions found; how many times heavier than hydrogen is it?

**90a.** Summarize your work upon ammonia, nitrogen, and the atmosphere.

It has been stated that air is a mixture of oxygen and nitrogen, and not a compound. (What is the difference?) There are, however, well-defined compounds of nitrogen and oxygen, and some of these will now be studied.

### **NITROUS OXIDE (Nitrogen Monoxide or Laughing Gas).**

**91.** Arrange a flask or retort of about 250 cc. capacity, as in the preparation of oxygen; place in it about 15 grams of ammonium nitrate, and heat gently; after the air is expelled from the apparatus, collect three cylinders of the gas over warm water. Why use warm water? This is nitrous oxide gas. Note its odor, color, and taste. Does it burn? Does it support combustion? Bring a lighted candle or splinter into one of the cylinders containing the gas.

**92.** Bring into a second cylinder of the gas a small piece of burning phosphorus in an iron combustion spoon; into the third cylinder bring a small piece of burning sulphur in an iron combustion spoon.

What gas does nitrogen monoxide resemble?

**93.** Take a glass tube about 20 cm. long and 7 or 8 mm. internal diameter. Seal one end and bend it near the middle almost to a right angle, as shown in Fig. 30. Place a small piece of potassium in the

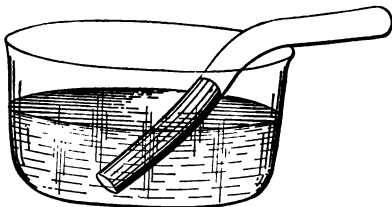


FIG. 30.

sealed end of the tube ; then fill with mercury, and invert over a mercury dish. Fill the tube about  $\frac{1}{2}$  full of dry nitrogen monoxide, marking carefully on the tube the volume after equalizing the pressure.

Then carefully heat the potassium. After all action has ceased, allow to cool, equalize the pressure, and measure the new volume.<sup>1</sup>

Close the mouth of the tube with the thumb, and invert it ; bring a lighted match into the remaining gas. What gas does it resemble? The potassium has taken out and united with all of the oxygen in the compound, leaving only the nitrogen. How does the volume of nitrogen compare with the original volume of gas?

Nitrogen monoxide has been found to be 22 times as heavy as hydrogen.

What volume of oxygen unites with one volume of nitrogen to form nitrogen monoxide, and what volume of nitrogen monoxide is produced?

Calculate the percentage composition of nitrogen monoxide.

### NITRIC OXIDE (Nitrogen Dioxide).

94. Take a flask of 250 cc. capacity provided with a funnel tube and a delivery tube. Place in it a quantity of copper turnings, and just cover them with water. Pour strong nitric acid through the funnel tube. Allow the action to go on until there are no more red fumes formed in the flask ; then collect five cylinders of the gas over water.

<sup>1</sup> This experiment must be performed with *great* care to get satisfactory results. It is indispensable that the mercury be perfectly dry and clean; that the nitrous oxide be pure and dry; and that the potassium be freed carefully from all adhering oil, and that its surface be freshly cut and as little tarnished as possible.

Note the odor and color of the gas. It is nitrogen dioxide. Does the gas burn? Observe the change when the gas is allowed to escape into the air.

95. Lower a piece of burning phosphorus into a cylinder of the gas; into a second cylinder put a piece of burning sulphur; and into a third bring a piece of burning wood or a lighted candle. Does it support combustion?

96. Open a cylinder of the gas to the air, and observe what takes place. Nitric oxide combines with the oxygen of the air to form a new compound called *nitrogen tetroxide* or *nitrogen peroxide*.

97. Into the fifth cylinder, with its mouth still below the surface of the water, pass a slow current of pure oxygen from a gasometer; observe carefully all the changes, and explain what you see.

98. Determine the amount of nitrogen there is in a known volume of nitric oxide exactly as you determined the amount of nitrogen in nitrous oxide in Exp. 93.

Nitric oxide, as prepared in Exp. 94, contains some other gases. To make the pure nitric oxide needed in this experiment, into a small flask or retort put 6.5 grams potassium nitrate, 65 grams ferrous sulphate, and 85 cc. dilute sulphuric acid (1 volume of acid to 3 volumes of water), and heat the mixture. Observe the same precautions as in Exp. 93.

Nitric oxide has been found to be 15 times as heavy as hydrogen. Calculate the percentage composition of this gas by volume and by weight.

What volume of oxygen unites with one liter of nitrogen, and how much nitric oxide would be formed?

**CARBON.**

**99.** Examine a piece of charcoal; note its structure and general physical properties. Is it soluble in water? In alcohol? In acids? Heat some of it to redness; what occurs? How is charcoal made?

**100.** Take about 50 cc. brown sugar solution, or starch colored with iodine, or red wine, or indigo solution, or litmus solution, or, in fact, almost any solution colored with vegetable or animal coloring matter, and place in it  $\frac{1}{2}$  teaspoonful of bone charcoal, and warm a few moments; shake it well and then filter.

**101.** Fill a test tube with mercury; invert it over a mercury bath, and fill the tube with dry ammonia gas, made by warming a strong solution of ammonia and drying the gas by passing over quicklime. Heat a piece of charcoal about the size of a marble, press it under the mercury while still hot, and allow it to rise to the surface under the tube full of ammonia gas.

This power of charcoal to absorb coloring matter, gases, etc., causes it to be used in purifying water in charcoal filters, in the refining of sugar to bleach it, and sometimes as a disinfectant to absorb noxious gases and bad odors.

Charcoal is an impure form of the element carbon.

**102.** Examine a piece of soft coal, heat some of it in a glass tube, and note what occurs. Collect some of the gas given off; note its odor; does it burn? What is it? When no more volatile matter is given off, examine the residue. It is called coke; like charcoal, it is an impure form of the element carbon.

**103.** Hold a piece of cold porcelain in the flame of a candle, or in the flame of a Bunsen burner with the air

supply cut off; it becomes coated with soot; this, too, is impure carbon.

**104.** Graphite (the graphite of lead pencils) is impure carbon. How does it differ from coke and charcoal?

The diamond is another form of the element carbon; how does it differ from graphite and charcoal?

The three substances, charcoal, graphite, and diamond, are each composed of the same element, carbon, yet they differ very markedly in their properties; this phenomenon is exhibited by a few other elements, and is known as *allotropism*.

Define allotropism.

**105.** What happens when charcoal burns in the air? In oxygen?

**106.** Arrange an apparatus as shown in Fig. 31. *A* contains a solution of caustic soda; *B* contains calcium chlo-

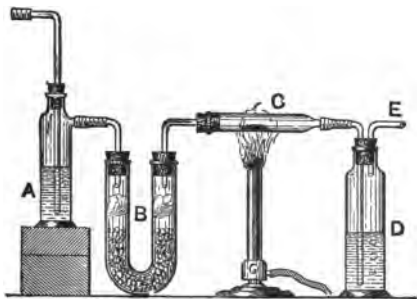


FIG. 31.

ride; *C*, several small pieces of charcoal; and *D*, some clear lime-water. The outlet tube *E* is attached to the filter-pump, or an aspirator, and a slow current of air drawn through the apparatus for several moments; is there any change in the lime-water? Finally, heat the charcoal

in *C* to redness until it slowly disappears. Note the change in the lime-water.

The compound formed when carbon burns is known as *carbon dioxide*, or *carbonic acid gas*.

**107.** Mix thoroughly 1 or 2 grams of powdered copper oxide with about  $\frac{1}{2}$  its weight of powdered charcoal, and heat in a test tube, first closing the tube with a cork and delivery tube, so as to pass any gas that may be formed into the lime-water. What remains in the test tube?

### CARBON DIOXIDE

**108.** Arrange an apparatus as shown in Fig. 32; in the flask *A* are small pieces of marble; the wash bottle *B* contains strong sulphuric acid. Pour into the funnel tube dilute hydrochloric acid; after the apparatus has run for some time, collect two cylinders full of the gas by downward displacement, as shown in the figure. This gas is carbon dioxide; note its color and odor. Is it heavier than air? Where did the gas come from? Does the gas burn? Try to ignite the gas as it comes from the delivery tube. Does it support combustion? Lower into the gas a burning candle.

**109.** Pour the second cylinder of gas over a burning match; pour it out of the jar just as though it were water.

**110.** Fill a test tube with carbon dioxide over mercury; then, covering the tube with the thumb, transfer it, mouth downward, to a vessel of water, removing the thumb after the mouth of the tube is below the surface of the water. Allow it to stand for some time. Does the gas dissolve in water?

**111.** Add some caustic soda to the water, and then repeat the previous experiment. Does the gas dissolve in caustic soda solution?



**112.** Pass the gas into a test tube containing distilled water ; after some moments, test the water with blue litmus paper. Taste the water.

**113.** Pass the gas for 10 or 15 minutes into a test tube containing a caustic soda solution ; after this, add hydrochloric acid to the solution. What gas results? Explain what has taken place.

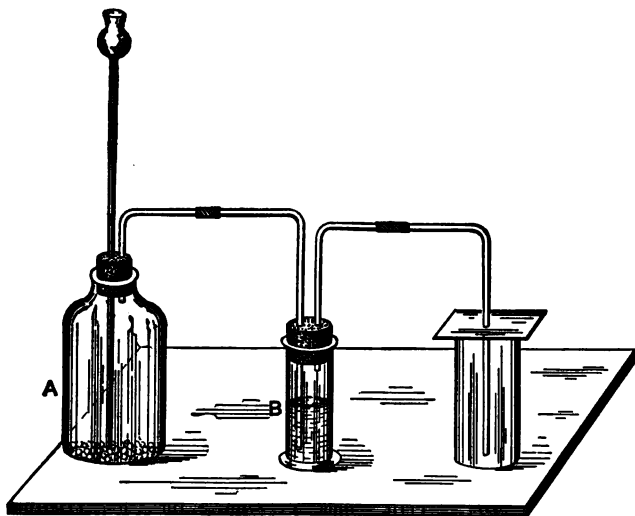


FIG. 32.

**114.** Pass some of the gas into lime-water ; filter out the white substance, and treat it with hydrochloric acid. How does it act? What is it?

**115.** Blow your breath through a solution of lime-water. What conclusions can you draw from this experiment?

Besides carbon dioxide, carbon and oxygen form another compound containing only half as much oxygen as carbon dioxide ; it is called *carbon monoxide*, or *carbonic oxide*.

**CARBON MONOXIDE**

**116.** Arrange a flask of 300 to 500 cc. capacity, with a funnel tube and a delivery tube. Place in it 50 grams of crystallized oxalic acid; pour 100 cc. of concentrated sulphuric acid over it, and heat gently. A mixture of gases is given off; pass this mixture into lime-water. Do you recognize either of the gases?

Pass the gases through a strong solution of caustic soda. What does this do? Collect some of the gas in a test tube. Note its color. The gas is poisonous, and should be handled with care. Pass some of the gas that passes through the caustic soda solution into lime-water.

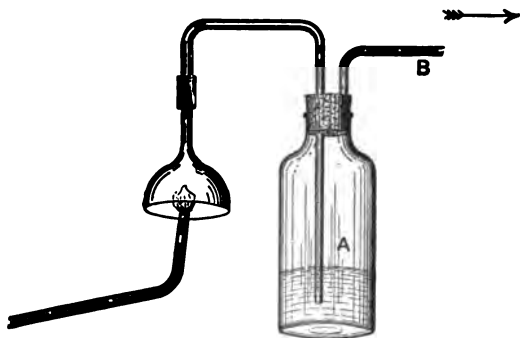


FIG. 33.

**117.** Ignite the gas as it comes from the generator after passing through the caustic soda solution; bring the burning jet under a funnel arranged as shown in Fig. 33, so as to draw the product of combustion through the lime-water in *A*. Connect *B* with the air pump or aspirator.

**118.** Pass the gas through water, and test with litmus, as in the case of carbon dioxide; also through a solution of

caustic soda ; and finally treat this solution with acid, as in the case of carbon dioxide. What are the differences?

**119.** Place some copper oxide in a straight, hard, glass tube, like that shown in Fig. 8 ; pass some carbon monoxide over the heated copper oxide, and then through lime-water. Explain the changes that have taken place. Finally draw air over the red-hot copper oxide remaining in the tube. Explain the changes. Save the copper oxide.

**120.** Place some pieces of charcoal in the tube used in the previous experiment ; connect one end of the tube with an apparatus generating carbon dioxide, and the other with a bottle containing caustic soda. Heat the charcoal to redness ; collect over water some of the gas which passes through the caustic soda, and examine it. What is it? Explain the changes that have taken place.

### **METHANE (Marsh Gas).**

**121.** Heat about 15 grams of sodium acetate upon an iron plate until it is dry, then mix it with about 20 grams of soda lime ; charge a retort, or flask, with the mixture ; heat gently, and collect two cylinders of the gas over water. Observe the color and odor of the gas. Does it support combustion? Does it burn? Is it acid or alkaline?

**122.** Light the stream of gas as it comes from the generator ; hold a beaker of cold water over the flame. What is deposited on the sides of the beaker?

Arrange the burning jet so that the products of combustion pass through lime-water. (See Fig. 33.) Of what is marsh gas composed?

**123.** Fill a test tube over water  $\frac{1}{3}$  full of marsh gas, and then fill the remainder of the tube with oxygen ; cover the mouth of the tube with the thumb ; invert the tube ; upon

removing the thumb bring a lighted match to the mouth of the tube.

Carbon and hydrogen unite with each other to form a great many other compounds called hydrocarbons.

### THE FLAME

A flame is produced by chemical changes between two gases; ordinarily it is the chemical action of oxygen upon some other gas. Is this true in the case of a burning candle, or burning wood?

Illuminating gases differ in their composition, but are usually made up of hydrogen, carbon monoxide, marsh gas, and other hydrocarbons, and contain also more or less oxygen, carbon dioxide, and nitrogen.

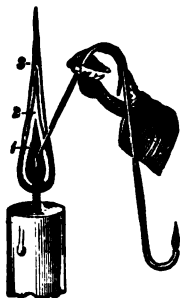


FIG. 34.

124. Insert the head of a match quickly into the inner cone of a Bunsen flame. What does this experiment show? Insert a bent glass delivery tube into the same part of the flame. Examine the gas that escapes. Does it burn? (See Fig. 34.)

Examine the gas taken from other parts of the flame in the same way. What would be the products of burning a gas whose composition was that given above? Prove to yourself that these are formed in the flame of the ordinary Bunsen burner.

Explain the changes that take place when the air holes at the base of the Bunsen burner are closed.

Make a drawing, illustrating the different parts of the flame.

125. As the flame is the result of chemical action between two gases, and illuminating gas burns in oxygen, it follows

that oxygen will burn in an atmosphere of illuminating gas. This may be shown by the following experiment. Take the neck of a broken retort, connect the small end with the gas jet, turn on the gas, and when the air is driven out of the retort neck, light the gas. (See Fig. 35.) Bring into the atmosphere of illuminating gas within the retort neck a jet of oxygen. The oxygen will be seen to burn. With a little care and patience a jet of air may be made to burn in the illuminating gas in the same way.

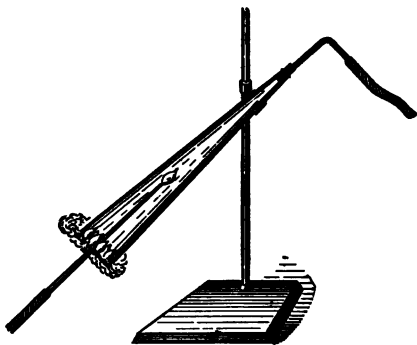


FIG. 35.

**126.** Light a Bunsen burner; bring down in the flame a piece of wire gauze; does unburned gas pass through? Try to light it above the gauze. Why does not the flame pass through? Hold the gauze for some time in the flame, and observe that after a while the unburned gas above the gauze ignites itself.

Turn on the gas, and, without lighting it, hold a wire gauze about 10 cm. above the burner. Light the gas above the gauze. Explain the phenomena observed.

The principle of the miner's safety lamp is included in these experiments.

## SULPHUR.

**127.** Examine some sulphur, and note its color, odor, and taste. Test its solubility in water, alcohol, acids, and carbon bisulphide. Sulphur is a chemical element.

**128.** Find the specific gravity of sulphur as follows: Procure a small flask of not more than 50 cc. capacity; fill it with distilled water, which has been previously boiled and then cooled, so that the water is within one or two centimeters of the top; carefully mark the level of the water in the neck of the flask with a label or scratch, and weigh carefully. Now weigh out about 5 grams of sulphur (note the exact weight of the sulphur you take); put it into the flask with the water; remove water from the flask with blotting paper until the former level is reached; weigh again, and from the data thus obtained, calculate the specific gravity of sulphur.

**129.** Heat some pieces of sulphur in a dry test tube, noting carefully the changes that take place as the temperature increases. When the sulphur begins to boil, remove the heat, and pour some of the molten sulphur into cold water. After a moment take it out of the water and examine it. What changes does it undergo upon standing for some time exposed to the air?

**130.** In a small retort place a small quantity of sulphur; heat the sulphur to boiling; collect any sulphur that distills over, by allowing it to drop into a beaker containing distilled water. Examine some of the sulphur powder that collects in the cooler portions of the neck of the retort. Is it soluble in carbon bisulphide? Examine the sulphur that has collected in the water; compare its properties with those it had before heating. Allow this distilled sulphur to stand exposed to the air for some time, and again examine, noting carefully the changes in its properties.

**131.** In a small retort, *A* (Fig. 36), place a quantity of sulphur, and connect the retort with the large flask *B*. Heat the sulphur in *A* until it boils. Explain the phenomena you observe. The powder collected in *B* is called *flowers of sulphur*.

**132.** In a small Hessian crucible place a quantity of sulphur, and heat it until it is quite fluid; then remove the

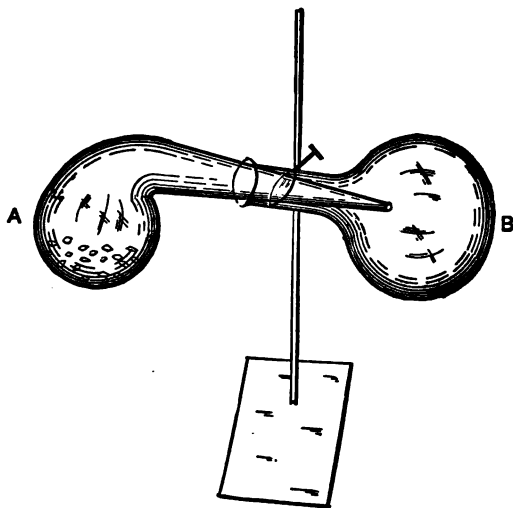


FIG. 36.

heat and allow it to stand until a crust forms on the surface of the molten sulphur. Cut the crust and pour out the liquid interior; examine the sulphur on the sides of the crucible, noticing carefully the shape of the crystals.

**133.** Dissolve some powdered sulphur in 2 or 3 cc. of carbon bisulphide, and allow the carbon bisulphide to evaporate spontaneously from a small evaporating dish. (Carbon

bisulphide is very inflammable ; have no lights near it.) Examine the crystals left ; have they the same shape as those in the previous experiment ?

**134.** Mix intimately about 7 grams of fine iron filings and 4 grams of powdered sulphur. Put the mixture into a test tube, and heat to redness ; finally, cool and break the tube, powder the residue, and examine it carefully. Is it magnetic ? Is it soluble in carbon bisulphide ? What has become of the iron and the sulphur ?

**135.** Put a bright piece of copper into some molten sulphur ; what changes does it undergo ?

We have seen that sulphur combines with oxygen at high temperatures, or, in other words, that sulphur will burn in oxygen or in the air. The product of this combustion is called *sulphur dioxide*.

### SULPHUR DIOXIDE

†**136.** Burn a little sulphur in the air, and note the color and odor of the product formed.

Put a small piece of sulphur on the lid of a porcelain crucible ; ignite it, and then bring it under a tripod or other support upon which are some fresh flowers of different colors ; and over the whole place a bell jar to confine the products of combustion. What changes do the flowers undergo ?

†**137.** Into a 500 cc. flask put some metallic copper in the form of strips or turnings ; close the flask with a rubber stopper carrying a funnel tube and a delivery tube. Pour through the funnel tube enough concentrated sulphuric acid to just cover the metallic copper ; then heat the flask, collecting cylinders of the gas by downward displacement of the air. Does the gas burn ? Does it support combustion ?



Test the gas with pieces of moist litmus paper. Is it acid or alkaline? Avoid breathing the gas.

†138. Invert one of the cylinders of gas over water and allow it to stand for a few moments. Is the gas soluble in water?

Pass a stream of gas for a moment through water in a test tube; examine the water. What are its properties?

†139. Take a little sodium hydroxide solution in a test tube; pass the gas through it for some time; finally, evaporate the solution to dryness, and add a few drops of hydrochloric acid to the residue. What happens?

The solution of sulphur dioxide in water is called *sulphurous acid*: and one of its most characteristic properties is to combine with oxygen and be converted into sulphuric acid. As it has the power of taking oxygen from other things, it is called a *reducing agent*.

†140. Pass a stream of sulphur dioxide through a solution of potassium permanganate to which a little sulphuric acid has been added. What change do you observe? Pass the gas through a solution of potassium bichromate to which a little sulphuric acid has been added. Observe the change.

141. Does a solution of the gas in water behave the same as the gas? Add a little of this solution, first to the potassium permanganate, and then to the potassium bichromate.

### **HYDROGEN SULPHIDE (Sulphuretted Hydrogen).**

†142. Take a 250 cc. flask; place in it a few pieces of iron sulphide, and close the flask with a stopper carrying a funnel tube and a delivery tube. Pour through the funnel tube some dilute sulphuric acid. Collect some of the gas, and answer the following questions: Is it soluble in water?

Is it acid or alkaline? Does it support combustion? Does it burn? What is its odor and color?

†143. What are the products of combustion? Observe the odor produced by the burning gas; do you recognize it? Hold a cold beaker over the flame. What conclusions can be drawn in regard to the composition of hydrogen sulphide?

†144. Pass the gas into a solution of potassium permanganate; again, into potassium bichromate, as in Exp. 140. What changes are observed?

†145. Pass a stream of this gas, together with a stream of sulphur dioxide, into a large flask.

†146. Take a test tube half full of each of the following solutions: copper sulphate, arsenic chloride, antimony chloride, lead acetate, and zinc acetate, and pass some of the gas into each tube.

147. Repeat Exp. 146, using, instead of the gas, some water through which the gas has been passed.

148. Into a bent tube of hard glass (Fig. 30) filled with mercury introduce dry sulphuretted hydrogen, noting carefully the volume of gas in the tube. Place a piece of tin in the bent portion of the tube, and heat it. Finally, allow to cool and measure the volume of the gas remaining. What is it? What has become of the tin?

### SULPHURIC ACID.

149. Pour 10 cc. of commercial concentrated sulphuric acid (called oil of vitriol in the trade) into 20 cc. of water. What is the effect? Save this dilute acid for other experiments.

150. Put some bits of paper and splinters of wood into some of the concentrated acid in a beaker.

**151.** Add, drop by drop, a little of the concentrated acid to a few cubic centimeters of concentrated sugar solution.

**152.** Take a little of the dilute acid formed in Exp. 149, and add to it a few drops of a solution of barium chloride.

**153.** What is the effect of this dilute acid on red and blue litmus paper?

**\*154.** The manufacture of sulphuric acid can be illustrated by means of the apparatus represented in Fig. 37. This

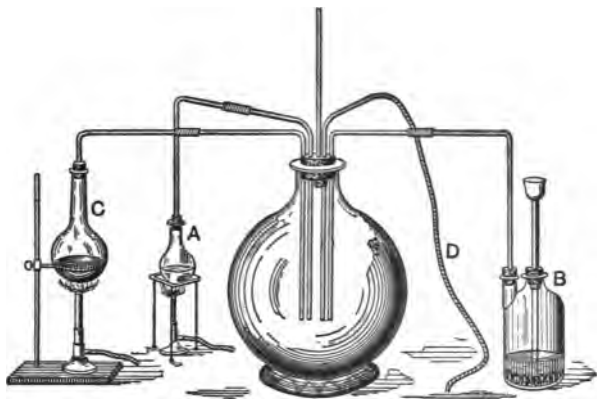


FIG. 37.

consists of a large 4 or 5 liter flask, fitted with a cork having five openings. By means of tubes it is connected with three small flasks. One of these, *A*, contains water, for the purpose of providing a current of steam; another, *B*, contains copper turnings, or foil, and dilute nitric acid, which give nitric oxide; and the third, *C*, contains copper foil and strong sulphuric acid, which give sulphur dioxide when heated. The flasks *A* and *C* are heated; steam, sulphur dioxide, and nitric oxide, are thus brought together in the flask; air is forced into a fourth opening *D* by means of

a foot bellows. These gases react upon one another, and produce sulphuric acid.

**155.** Review Exps. 23, 49, 50, 52, 58, 61, 64, 84, 116, 137, and 142, in each of which sulphuric acid has been used. Explain its use in each case. Make as complete a summary of the properties of sulphuric acid as possible.

### NITRIC ACID.

**156.** Arrange a 250 cc. retort and receiver so that a stream of cold water may be run over the receiver. (See Fig. 38.) Place in the retort 40 grams potassium nitrate, or, better,

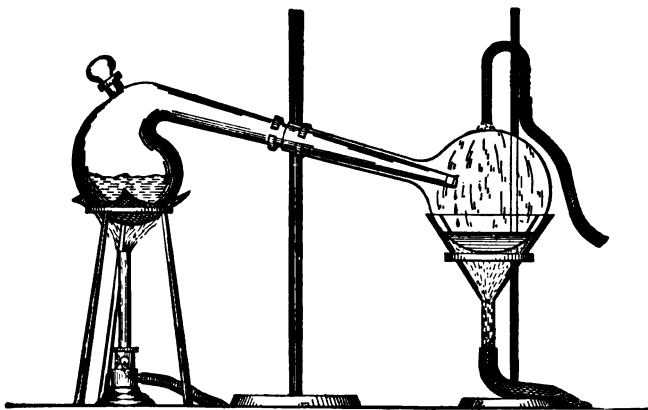


FIG. 38.

34 grams sodium nitrate, and 25 cc. concentrated sulphuric acid. Heat the retort gently as long as any liquid distills over. What is the odor and color of the liquid? What is its effect on litmus paper?

This liquid is nitric acid. When pure it is colorless, but

it usually contains some oxides of nitrogen, which give it a color. Compare the production of nitric acid in this experiment with the production of hydrochloric acid in Exp. 61.

**157.** Add a solution of sodium hydroxide to 10 cc. of the nitric acid made in the previous experiment, until the solution has no effect on litmus paper. Evaporate the solution, and compare the product obtained with sodium nitrate used in Exp. 156. What is the difference?

(If potassium nitrate was used in Exp. 156, use potassium hydroxide in this experiment.)

**158.** Take five test tubes, each containing about 5 cc. of concentrated nitric acid. Drop into one of these a piece of zinc; into another, a piece of copper; into another, a piece of lead; into the fourth, a piece of tin; into the fifth, an iron nail. Observe carefully the changes in each case; if, after five minutes, in any case, there is no apparent change, warm the test tube. Do you recognize any of the products formed?

**159.** Repeat Exp. 158, using, instead of concentrated acid, acid which has been diluted with three times its own volume of water. What differences do you observe? Do all the metals go into solution? Evaporate each solution at a gentle heat, until only 3 cc. of the solution remain, and allow it to cool. Examine the residue in each case. What is it? Is it soluble in water? How does heat affect it?

**160.** Add one cubic centimeter of concentrated nitric acid to 4 or 5 cc. of indigo solution, and warm.

**161.** In a small beaker, cover some powdered sulphur with the strongest nitric acid, and warm gently for some time. Finally, dilute with water, and add a few drops of barium chloride. Explain what has occurred.

**162.** Pass a stream of sulphuretted hydrogen into a test tube containing nitric acid. Explain the results.

**163.** Mix two volumes (about 10 cc.) of strong hydrochloric acid with one volume (about 5 cc.) of strong nitric acid. What gas is given off? How was the gas obtained before? Explain how it comes to be given off here.

### ACIDS AND BASES.

**164.** Review what you have learned with regard to hydrochloric acid, sulphuric acid, and nitric acid. What properties have they in common?

What do we mean when we say that a substance has acid properties?

There are many acids known to chemists; the three already studied are the most common.

**165.** Take 25 cc. of concentrated sulphuric acid, and put it into 150 cc. of distilled water; into a small beaker measure 10 cc. of this diluted acid; make 500 cc. of a solution of caustic soda, and allow it to run into the acid from a burette, until the solution, upon stirring, has no influence on litmus paper. If a drop of phenol-phthalein solution is put into the acid, and the caustic soda allowed to flow slowly, and finally drop by drop, into the acid, until a faint pinkish color is observed, the neutral point can be obtained with greater accuracy and less trouble. Repeat the experiment, using 5 cc. of acid instead of 10 cc. Note accurately the amount of caustic soda required to neutralize the acid in each case. Do these amounts bear the same ratio to one another as the amounts of acid taken?

**166.** Evaporate to dryness the solution left in the previous experiment and examine it. Has it acid properties? Has it the properties of caustic soda? What is it?

**167.** Repeat Exps. 165 and 166, using caustic potash instead of caustic soda.

**168.** Repeat Exps. 165 and 166, using a solution of ammonia instead of caustic soda.

**169.** Repeat Exps. 165, 166, 167, and 168, using nitric acid instead of sulphuric acid.

**170.** Repeat Exp. 169, using hydrochloric acid instead of nitric acid.

Substances which possess the power of destroying the properties of an acid, such as caustic soda and caustic potash, are called *bases*. Bases and acids are two classes of substances which act upon one another, each destroying the properties of the other, forming a neutral substance; *i.e.* a substance that will not act upon litmus and other vegetable colors,—a substance having neither acid nor basic properties. These neutral substances are called *salts*. The salts formed by neutralizing sulphuric acid are called *sulphates*; those formed from nitric acid are called *nitrates*; and those derived from hydrochloric acid are called *chlorides*.

### THE METALLIC ELEMENTS.

**171.** What are the distinctions between the metallic elements and the non-metallic elements?

#### Sodium.

**172.** Cut off a few thin chips of metallic sodium, and leave them exposed to the air. Observe the changes. Review Exp. 41. Is sodium lighter or heavier than water? Give briefly the properties of metallic sodium.

The more common compounds of sodium are the chloride,  $\text{NaCl}$ ; the sulphate,  $\text{Na}_2\text{SO}_4$ ; the nitrate,  $\text{NaNO}_3$ ; the carbonate,  $\text{Na}_2\text{CO}_3$ ; and the hydroxide,  $\text{NaOH}$ : all of which have already been met with. Review Exps. 12,

13, 15, 16, 61, 63, 64, 76, 113, 156, 157, 166, 169, and 170.

All sodium salts are soluble in water.

**173.** Make a solution of sodium chloride (common salt) and sodium nitrate or any other sodium salt; heat a piece of platinum wire in the flame; then dip it into the salt solution and heat it again. What effect does it have on the flame? Do the solutions of the other salts of sodium have the same effect? Look at the sodium flame through a piece of blue glass, what effect has the blue glass on the appearance of the flame?

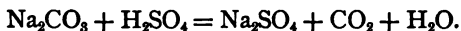
**174.** Dissolve 20 grams of sodium carbonate in 100 cc. of hot distilled water, and filter if necessary. Evaporate the solution down to about 75 cc., and allow to cool. Separate the crystals from the mother liquor, and examine them carefully. Do they contain water? Take about half of these, and heat them nearly to redness in a porcelain crucible. Keep these for use in the next experiment.

**175.** Prepare 250 cc. of dilute sulphuric acid, and find how much sulphuric acid there is in one cubic centimeter of the dilute acid, as follows:—

Weigh out accurately about one gram of the ignited sodium carbonate, made in the previous experiment; place in a porcelain dish; dissolve in a little water; add two or three drops of phenol-phthaleïn; heat the solution to boiling, and allow the dilute acid to run into it from a burette, until the liquid remains colorless after continued boiling. The carbonate is then exactly neutralized. The phenol-phthaleïn takes no part in the reaction; it merely indicates, by its change of color, the complete neutralization of the alkali. The addition of acid should be stopped the instant the color changes, and the amount of acid added should be read off. From these data calculate the amount of sulphuric acid in



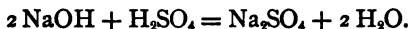
one cubic centimeter of your dilute acid. The change is represented by the equation —



Keep this standard acid in a well-stoppered bottle for future experiments.

**176.** Find the amount of caustic soda in a dilute solution of caustic soda as follows : —

Take 10 cc. of dilute sodium hydroxide solution in a porcelain dish. Add a drop of phenol-phthaleïn, and then run the standard acid from a burette into the solution, until the color just disappears. Knowing the strength of the acid, calculate the amount of sodium hydroxide in each cubic centimeter of the solution from the equation —



### Potassium.

**177.** Cut off a very thin piece of metallic potassium, noting the appearance. Allow it to remain exposed to the air, and observe the changes it undergoes. Review Exp. 42. Give briefly the properties of metallic potassium, and compare them with the properties of metallic sodium.

The more common compounds of potassium are similar to the ones given for sodium. Give their names and symbols.

**178.** Repeat Exp. 173, using potassium salts instead of sodium salts. What is the difference in the appearance of the flames? What effect has blue glass upon the appearance of the flame? Make a mixture containing both potassium and sodium chloride. What effect has this mixture upon the flame when seen through blue glass? What, when seen without the glass? Could you distinguish between potassium and sodium compounds in this way?

**179.** Weigh a small watch glass with care; then weigh out quickly three or four very small pieces of metallic potassium, not more than half a gram in all. Measure out 75 cc. of distilled water in a tall beaker; into this water drop the pieces of potassium, one at a time, covering the beaker each time with a glass plate, until action is over. After all the weighed potassium is added, put in a drop of phenolphthalein. Place the beaker on a piece of white paper, and run into it from a burette the standard acid made in Exp. 175 until the color just disappears. Calculate then how much acid it requires to neutralize the potassium used, and from these data calculate the combining weight of potassium, knowing that the equation is —



### **Ammonium Compounds.**

**180.** Review your experiments on ammonia, Exps. 77 to 84.

Purify some commercial ammonium chloride by dissolving 20 grams of it in 150 or 200 cc. of distilled water, and add to this solution a little solution of the gas in water (ammonium hydroxide). If a precipitate is formed, filter it out, evaporate the solution to small bulk, and allow it to crystallize. Separate the crystals from the mother liquor. Compare these crystals with the unpurified crystals. Heat some of the crystals in a porcelain dish. Do they color the flame?

**181.** Dissolve some of the crystals of ammonium chloride in water in a test tube. Add a little caustic soda solution, and warm. Do you recognize the odor? Bring a moist piece of red litmus paper to the mouth of the tube. Bring a glass rod moistened with hydrochloric acid to the mouth

of the tube. All salts of ammonium behave in this way, so that their presence in a substance can be readily detected.

### Calcium Compounds.

The most common compounds of calcium are calcite, aragonite, and marble, which are nearly pure calcium carbonate, and limestone, a more impure variety of the same compound.

**182.** Heat 2 or 3 grams of powdered marble in a porcelain crucible to the highest heat obtainable with a blast lamp for 20 minutes, stirring the powder from time to time. The marble,  $\text{CaCO}_3$ , is converted into quicklime,  $\text{CaO}$ . After cooling the lime, add a little water; finally add more, 200 cc. in all; shake thoroughly, and allow to settle. When clear, pour off. This is lime-water.

Review Exps. 114 and 115.

Lime-water is a solution of calcium hydroxide,  $\text{Ca}(\text{OH})_2$ .

Add a little lime-water to a solution of ferric chloride. Write the equations for all the changes in this experiment.

**183.** Obtain a little powdered gypsum (calcium sulphate). Does it contain water of crystallization? Dry it thoroughly, avoiding a heat greater than  $200^\circ$ . Mix some of the dried gypsum into a paste with water. What changes does it undergo on standing? Dry it again, powder it, and moisten again with water. Does it behave in the same way? This dried gypsum is known as plaster of Paris. Can you give any reason for its peculiar action?

**184.** Give some account of calcium chloride. What have you used it for in previous experiments? Dissolve a little in water. What color does it give to the flame? Try it with a platinum wire. Add a little sodium carbonate solu-

tion to some of your calcium chloride solution. Add a little sulphuric acid to some of the calcium chloride solution. Explain results. What changes have taken place?

### **Zinc and its Compounds.**

**185.** Examine some metallic zinc. What are its properties? Find its specific gravity by the method used for sulphur (Exp. 128). What is the effect upon it of hydrochloric acid? Of sulphuric acid? Of nitric acid? Write the equation in each case.

Heat a small piece of the metal on a piece of charcoal with a blowpipe flame. Does the metal burn? What is the coating formed on the coal? Notice its color while hot and when cold. Moisten with a drop of cobalt nitrate, and then heat again. What is the effect? What is the equivalent weight of zinc? Review Exps. 50 and 51.

**186.** Obtain some crystals of zinc sulphate. How can zinc sulphate be made? Is it soluble in water? Has it water of crystallization?

Add a solution of sodium hydroxide,  $\text{NaOH}$ , to zinc sulphate solution. What is formed? What is the effect of adding more sodium hydroxide? Add a solution of sodium carbonate,  $\text{Na}_2\text{CO}_3$ . What is formed? Add ammonium sulphide,  $(\text{NH}_4)_2\text{S}$ . What is formed?

### **Copper and its Compounds.**

**187.** Obtain a strip of metallic copper, and examine it. Is copper specifically lighter or heavier than zinc? Try the effect of the three acids upon it in the concentrated and in the dilute form, also cold and hot. What is formed in each case? Give the equations.

**188.** Obtain some crystals of copper sulphate. How can they be made? Are they soluble in water? Have they water of crystallization? How much? (See Exps. 6, 9, 11, and 12.) What is the effect of passing a current of electricity through a solution of copper sulphate?

Add a solution of sodium hydroxide, of sodium carbonate, of ammonium sulphide, respectively, to different samples of the copper sulphate solution.

**189.** To find the equivalent weight of copper, dissolve about 10 grams of copper sulphate in 100 cc. of water. Filter if the solution is not perfectly clear. Weigh out exactly one gram of pure zinc. Put the zinc in the copper sulphate solution; cover the beaker carefully with filter paper to keep out dust, and set aside for 24 hours. What has taken place at the end of this time? Filter out the copper that has been thrown out of the solution, carefully wash it with distilled water, dry it at a temperature not above 100°, transfer it to a watch glass, and weigh it carefully. This amount of copper is equivalent to one gram of zinc. What reason have we for assuming this? What chemical change has taken place here?

### **Silver and its Compounds.**

**190.** Dissolve a ten-cent piece in a little dilute nitric acid in a small beaker. After all action is over, evaporate to dryness in a water bath, and then dissolve the residue in 100 cc. of distilled water. The color is due to the presence of copper nitrate along with the colorless silver nitrate. All United States silver coins contain about 10 per cent of copper. Divide this silver solution into two parts, one of 25 cc., and the other of 75 cc., for the next experiment. To the 25 cc. add dilute hydrochloric acid, drop by drop,

until no more precipitate is formed. Filter out the precipitate. This is silver chloride,  $\text{AgCl}$ . The copper remains in solution, and is thus separated from the silver. Heat some of this silver chloride on charcoal in the reducing flame of the blowpipe. Little globules of metallic silver are formed. These, carefully freed from adhering dirt, may be again dissolved in dilute nitric acid, evaporated in a water bath, and finally dissolved in distilled water, thus forming a solution of silver nitrate,  $\text{AgNO}_3$ , free from copper. Test the behavior of different portions of this solution with sodium hydroxide, sodium carbonate, sulphuretted hydrogen, etc.

**191.** To find the equivalent weight of silver, weigh out 0.3 of a gram of clean metallic copper. Put it into the 75 cc. of silver solution reserved from the previous experiment, and allow to remain 24 hours. Filter out the precipitate of silver, wash, dry, and weigh. Calculate the weight of silver equivalent to one gram of copper. Knowing the equivalent weight of copper from Exp. 189, what is the equivalent weight of silver?

### Iron and its Compounds.

**192.** Dissolve two or three iron nails in 100 cc. hydrochloric acid, with the aid of heat. After the action is over, take about one cubic centimeter of the solution, and add ammonia solution to it. Observe the color and nature of the precipitate. Take half the remaining solution, add to it 3 cc. strong nitric acid and 10 cc. strong hydrochloric acid, and boil, carefully observing changes. Add a little ammonia solution to one cubic centimeter of this boiled iron solution, noting carefully the differences. Take some of the same solution in a test tube, and pass sulphur dioxide gas through it for 20 minutes, and then add a little ammonia to the solution. How does this product compare with the first?

193. Determine the equivalent weight of iron by repeating Exp. 189, using one gram of iron instead of zinc, the equivalent weight of copper being known from Exp. 189.

Iron forms two classes of compounds, one in which its valence is two, and the other in which its valence is three. The former are known as *ferrous* compounds, and the latter as *ferric* compounds. In the first solution above, there is present ferrous chloride,  $\text{FeCl}_2$ ; in the second, there is present ferric chloride,  $\text{FeCl}_3$ . Ammonia added to the first solution gave ferrous hydroxide,  $\text{Fe}(\text{OH})_2$ ; and to the second, gave ferric hydroxide,  $\text{Fe}(\text{OH})_3$ .

Oxygen, chlorine, bromine, nitric acid, and other *oxidizing* agents convert ferrous compounds into ferric compounds. Nascent hydrogen, sulphuretted hydrogen, sulphur dioxide, and other *reducing* agents convert ferric compounds into ferrous compounds. Other metals besides iron form more than one class of salts; for example, mercury forms *mercurous* salts, in which the valence of mercury appears to be one and *mercuric* salts, in which the valence of mercury appears to be two. The chlorides are  $\text{HgCl}$  and  $\text{HgCl}_2$ , respectively. Tin forms stannous salts, in which the valence is two ( $\text{SnCl}_2$ ); and stannic salts, in which the valence is four ( $\text{SnCl}_4$ ).

Many other metals form more than one class of compounds. In general, oxidizing agents convert those compounds in which the valence is lower into those in which the valence is higher, and reducing agents act in the reverse way.

The metallic elements yield the following classes of derivatives: —

- (a) Compounds with oxygen, and with oxygen and hydrogen; or, *oxides* and *hydroxides*.
- (b) Compounds with chlorine, bromine, and iodine; or, *chlorides*, *bromides*, and *iodides*.

- (c) Compounds with sulphur, and with sulphur and hydrogen; or, *sulphides* and *hydrosulphides*.
- (d) Compounds with nitric and nitrous acid; or, *nitrates* and *nitrites*.
- (e) Compounds with sulphuric and sulphurous acids; or, *sulphates* and *sulphites*.
- (f) Compounds with chloric and chlorous acids; or, *chlorates* and *chlorites*.
- (g) Compounds with phosphoric and phosphorous acids; or, *phosphates* and *phosphites*.
- (h) Compounds with carbonic acid, or *carbonates*.
- (i) Compounds with boric acid, or *borates*.
- (j) Compounds with silicic acid, or *silicates*.
- (k) Compounds with chromic acid, or *chromates*.

Other classes of compounds are known, but those mentioned are the most important.

Only a few of these most important classes of compounds will now be studied.

### Oxides and Hydroxides.

**194.** How are oxides formed? Review Exps. 33, 35, 36, 37, 87, 182, 207, 208, and 216, in which oxides have been formed.

**195.** How can sodium hydroxide be formed? Review Exp. 41. On the large scale it is made from sodium carbonate and lime.

**196.** Take a solution of copper sulphate, and add to it a solution of sodium hydroxide. Copper hydroxide,  $\text{Cu}(\text{OH})_2$ , and sodium sulphate are formed. Write the reaction. Is copper hydroxide soluble in water? Most of the metallic hydroxides are insoluble in water, and all such insoluble hydroxides can be made in a manner similar to copper hydroxide.

**197.** Take the copper hydroxide obtained in the previous experiment, and boil it for several minutes with distilled water. Observe the changes. The copper hydroxide loses



water and is converted into the oxide,  $\text{CuO}$ . Express the change by an equation.

**198.** Take a solution of mercuric chloride,  $\text{HgCl}_2$ , and add a little solution of sodium hydroxide. The compound formed in this case is mercuric oxide,  $\text{HgO}$ ; express the changes by an equation.

### Chlorides.

**199.** Review Exps. 43, 48, 63, 71, 76, 79, 80, 82, and 170, in which chlorides are formed.

**200.** Add a little dilute hydrochloric acid to the copper hydroxide obtained in Exp. 196, and warm until it goes into solution. Copper chloride,  $\text{CuCl}_2$ , is formed. Evaporate the solution and obtain the copper chloride. Is it soluble in water? Most, but not all, metallic hydroxides dissolve in hydrochloric acid to form chlorides.

**201.** Take a little lead nitrate,  $\text{Pb}(\text{NO}_3)_2$ , solution in a test tube and add a few drops of hydrochloric acid. To another test tube containing a solution of lead nitrate or lead acetate,  $\text{Pb}(\text{C}_2\text{H}_3\text{O}_2)_2$ , add a little of a solution of common salt. Lead chloride is obtained in both cases. Is it soluble in cold water? In hot water? Would lead hydroxide be soluble in hydrochloric acid as copper hydroxide is?

**202.** Heat samples of sodium chloride and lead chloride in test tubes; do they suffer any change?

**203.** What happens to chlorides when they are heated with sulphuric acid?

### Sulphides.

**204.** Review Exps. 134, 135, 146, 147, in which sulphides are formed. How may sulphides be formed? What is the effect of treating sulphides with hydrochloric or sulphuric acid?

**205.** Take 100 cc. of strong ammonia solution and divide it into two equal parts; pass sulphuretted hydrogen into one of these parts until no more is absorbed. The ammonia solution is thus converted into ammonium hydrosulphide,  $\text{NH}_4\text{SH}$ . Then add the other part of the ammonia solution; this forms ammonium sulphide,  $(\text{NH}_4)_2\text{S}$ . Is it soluble in water? Add a little hydrochloric or sulphuric acid; what effect does this have? Add some of this ammonium sulphide to solutions of ferric chloride,  $\text{FeCl}_3$ ; zinc sulphate,  $\text{ZnSO}_4$ ; manganese sulphate,  $\text{MnSO}_4$ ; copper sulphate,  $\text{CuSO}_4$ ; cadmium chloride,  $\text{CdCl}_2$ . Sulphides of the metals are formed in every case. Express the chemical changes by equations.

#### Nitrates.

**206.** Review Exps. 94, 157, 158, 159, and 169, in which nitrates have been formed. What methods are available for the preparation of nitrates?

**207.** Make some copper nitrate by one of the methods; evaporate the solution to a small bulk, but not to dryness, and allow the copper nitrate to crystallize. How do these crystals compare with the crystals of copper sulphate? Do they contain water of crystallization? Heat some of them in a test tube. What gases are given off? What is the substance left behind?

**208.** Heat some crystals of lead nitrate. What gas is given off? What is left behind in the tube? Collect some of the gas over water if need be in order to examine it carefully?

**209.** Heat some crystals of sodium nitrate, and examine the gas given off. Is the sodium nitrate changed? To answer this question, after the heated nitrate has cooled, add some strong sulphuric acid to it, and warm; for com-

parison, add some strong sulphuric acid to the unheated nitrate. Do they behave in the same way?

**210.** What is the effect, in general, of heating nitrates? Of heating them with sulphuric acid?

### **Sulphates.**

**211.** Review Exps. 49, 50, 52, 64, 82, 152, 165, 166, 167, and 168, in which sulphates have been formed. What methods are available for the preparation of sulphates?

**212.** Prepare dilute solutions of barium chloride and lead nitrate, and a strong solution of calcium chloride; to a small quantity of each in a test tube, add a little sulphuric acid. Sulphates of the corresponding metals are formed in each case. If to the solutions of the salts any soluble sulphate is added instead of sulphuric acid, the same insoluble sulphates will be formed. Make solutions of small quantities of sodium sulphate,  $\text{Na}_2\text{SO}_4$ ; copper sulphate,  $\text{CuSO}_4$ ; and iron sulphate,  $\text{FeSO}_4$ ; and add them successively to solutions of barium chloride,  $\text{BaCl}_2$ ; lead nitrate,  $\text{Pb}(\text{NO}_3)_2$ ; and calcium chloride,  $\text{CaCl}_2$ . Write the equations expressing the changes taking place in each case.

**213.** Heat a sample of iron sulphate and a sample of sodium sulphate to redness in porcelain crucibles. What difference in behavior do you observe?

### **Carbonates.**

**214.** Review Exps. 113 and 114, in which carbonates have been formed.

**215.** Prepare sodium carbonate from sodium chloride, as follows: Pass carbon dioxide into a strong solution of

ammonia (about 100 cc.) until it is no longer absorbed. A solution of acid ammonium carbonate or ammonium bicarbonate,  $(\text{NH}_4)\text{HCO}_3$ , is thus obtained. Add this to a strong solution of sodium chloride, as long as a precipitate is formed. Filter off the precipitate and dry it, by spreading it upon layers of filter paper. Heat some of the salt when dry, and determine what gas is given off. When gas is no longer given off, allow to cool, and examine the residue.

It is sodium carbonate.

This process for making sodium carbonate from ordinary salt is known as the *Solvay* or *ammonia* process, and is used on the large scale.

**216.** Prepare solutions of small amounts of calcium chloride, barium chloride, ferrous sulphate, copper sulphate, lead nitrate; add to each a little solution of sodium carbonate. Filter out the precipitate in each case; take part of each precipitate, and heat it to redness in the lid of a porcelain crucible. What gas is given off? What is left in the crucible lid? Treat other parts of the different precipitates with different acids. What gas is given off? What product is formed in each case? Express all the chemical changes by equations.

**217.** Summarize the information you have gained in the last twenty-two experiments. State how each class of compounds may be formed, and any general properties of each class that you may have learned.

By a careful study of the color and relative solubility of the various compounds of the metals in water, acids, alkalies, alcohol, etc., chemists are able to prove their presence in any mixture, and to separate them from one another; this is the province of *Analytical Chemistry*. A more thorough and systematic study of the metals and their compounds is taken up under the head of *Qualitative Analysis*.

## PART II.



### CHAPTER I.

PHYSICAL CHANGE. — CHEMICAL CHANGE. — CHEMISTRY. — MEASUREMENT OF GASES. — ABSOLUTE TEMPERATURE. — QUESTIONS IN REVIEW.

ALL the changes going on in the material of the universe about us may be divided into two classes, — *physical* changes and *chemical* changes.

**Physical Change.** — Physical changes are those changes in the form and properties of matter which do not alter the essential nature of the material, do not change the composition of the substance.

For example, a piece of iron may be worked into many different forms ; it may be hot or it may be cold ; it may conduct a current of electricity, or it may be magnetized : in all of which conditions it will have different properties, but so long as it remains *iron*, the changes are physical. Again, wood may suffer a great variety of changes in form and other properties, but so long as it remains wood, the changes are physical.

**Chemical Change.** — A chemical change, on the other hand, is a change in the essential nature of the material, a change in its composition as well as a change in its properties.

For example, when iron rusts, the rust formed is a very different material from iron ; the product formed is no longer

iron; the change is a *chemical* change. When wood is burned, the product is no longer wood; the change is a chemical one.

Other examples of chemical changes going on about us in nature are the decay of animal and vegetable substances; the conversion of the sugar in fruit juices into alcohol, by fermentation; the souring of milk; and the burning of fuel in our stoves. Such processes, by which the composition, taste, smell, action, and other properties of substances are altered so that new bodies with quite different properties are formed from the old, are called *chemical* processes, and are due to *chemical* action or change.

**Chemistry.**—The study of chemical changes, their causes, and their results, and the laws governing them constitutes the science of *Chemistry*.

Chemistry is, however, not set off by any hard and fixed boundary from other natural sciences; it is closely related to all of them. It passes by imperceptible degrees into physics on one side, into physiology and biology on another, and into mineralogy and geology, on still another side: thus we have the names *physical* chemistry, *physiological* chemistry, *mineral* chemistry, and so on, to designate these intermediate fields of science. The measurement of gases and the tension of vapors belong perhaps to physics rather than to chemistry, yet a knowledge of these things is as indispensable to the chemist as to the physicist. The subject of crystallography mentioned in Chapter III. is usually considered under the head of mineralogy, but the knowledge of crystals is almost if not quite as useful to the chemist as to the mineralogist. The field of nature is a continuous whole; and however we may attempt to classify and subdivide it for our own convenience, we will find no natural division lines. All boundary lines are of our

creation, and we continually find nature utterly ignoring them.

**Measurement of the Volume of a Gas.**—The fact that the volume of a gas is changed by both temperature and pressure renders it necessary that gases be measured under the *same* conditions of temperature and pressure in order that there shall be a constant relation between their volume and their mass; *i.e.* in order that one liter of a gas shall always have the same weight, or that one gram of a gas shall always have the same volume. But it is not convenient to have the conditions of temperature and pressure always the same; they are continually changing from day to day. We have seen from Exp. 1 that if we know the volume at one pressure, we can tell the volume at any other pressure; likewise, from Exp. 2 we learn that if we know the volume at any temperature we can calculate the volume at any other temperature. For the sake of comparison, chemists and physicists have agreed to reduce all their volume measurements to the volumes that the gases would occupy if the pressure were 760 mm. and the temperature were  $0^{\circ}$ ; these have, therefore, been called *standard conditions*, and the process of calculation is frequently referred to as *reducing to standard conditions*.

It is customary to measure the pressure upon a gas in terms of the height of a column of mercury. When the pressure is given as 760 mm., it is meant that the pressure is equal to that exerted by a column of mercury 760 mm. high. A gas measured in an open vessel or in a vessel over mercury, or over water in which the level of the liquid is the same inside and outside the vessel, is measured under the atmospheric pressure: this pressure is determined by reading the barometer. As the pressure and temperature vary, it is necessary to read the barometer and thermometer

whenever a gas is measured, in order to reduce the volume to standard conditions. If the gas is measured in a tube over mercury or water, and the level of the liquid inside the

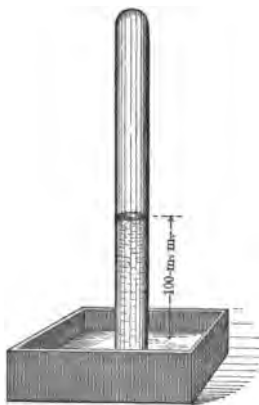


FIG. 39.

tube is higher than that outside, the gas is under diminished pressure, the amount of diminution depending upon the height of the column of mercury or water in the tube. Thus, if the arrangement be as represented in Fig. 39, and the height of the mercury column above the level of the mercury in the trough be 100 mm., and the pressure of the atmosphere be 755 mm., then the gas in the tube is not under the full atmospheric pressure, for the atmospheric pressure is supporting a column of mercury 100 mm. high besides the pressure

it is exerting upon the gas, so that the pressure actually brought to bear on the gas is

$$755 - 100 = 655 \text{ mm.}$$

If it were water in the tube instead of mercury, the pressure exerted by the water in terms of mercury must be determined. Since mercury is 13.6 times as heavy as water, the height of a column of mercury which would exert a pressure equal to 100 mm. of water would be  $\frac{100}{13.6}$  mm. Therefore, the pressure, in this case, would be

$$755 - \frac{100}{13.6} = 747.65 \text{ mm.}$$



Let  $V$  be the volume of a gas at a pressure of 760 mm., and  $v$  the volume at some other pressure  $p$ . From Boyle's law, that the volume of a gas varies inversely as the pressure it supports, we have —

$$V:v::p:760, \text{ or } V = \frac{vp}{760}. \quad (1)$$

This is a general expression for reducing the volumes of gases to standard conditions of pressure.

The law of Charles declares that all gases have the same coefficient of expansion by heat. Careful measurement has shown that a gas expands  $\frac{1}{273}$  (or 0.00366) part of its volume at  $0^\circ$ , for each degree the temperature is raised.

Hence it follows : —

$$\text{volume at } 1^\circ = \text{volume at } 0^\circ + \frac{1}{273} \text{ volume at } 0^\circ;$$

$$\text{volume at } 2^\circ = \text{volume at } 0^\circ + \frac{2}{273} \text{ volume at } 0^\circ;$$

$$\text{volume at } 3^\circ = \text{volume at } 0^\circ + \frac{3}{273} \text{ volume at } 0^\circ;$$

$$\text{volume at } t^\circ = \text{volume at } 0^\circ + \frac{t}{273} \text{ volume at } 0^\circ.$$

Let  $V$  be the volume at  $0^\circ$  and  $v$  be the volume at  $t^\circ$ ; then we have

$$v = V + \frac{t}{273} V = V \left( 1 + \frac{t}{273} \right) = V(1 + 0.00366 t).$$

$$\therefore V = \frac{v}{\left( 1 + \frac{t}{273} \right)} \quad \text{or} \quad V = \frac{v}{(1 + 0.00366 t)}. \quad (2)$$

This is a general expression for reducing the volumes of gases to standard conditions of temperature.

If  $V$  in formula (2) is to be the volume under standard conditions of both temperature and pressure, then  $v$  in this formula must be the volume under standard conditions of

pressure ; or, in other words, be the  $V$  of equation (1). The two expressions may be thrown into one by replacing  $v$  in (2) by the value of  $V$  in (1), and we get —

$$V = \frac{\frac{vp}{760}}{(1 + 0.00366 t)} = \frac{vp}{760(1 + 0.00366 t)}, \quad (3)$$

a general expression for reducing the volume of a gas to standard conditions where  $v$  is the volume at any temperature  $t^\circ$ , and any pressure  $p$ .

**Absolute Temperature.** — The law of Charles holds true for temperatures below zero as well as for temperatures above zero ; hence —

the volume at  $-1^\circ$  =  $\frac{273}{274}$  volume at  $0^\circ$  ;

the volume at  $-2^\circ$  =  $\frac{271}{274}$  volume at  $0^\circ$  ;

the volume at  $-3^\circ$  =  $\frac{270}{274}$  volume at  $0^\circ$  ;

the volume at  $-273^\circ$  =  $\frac{0}{274}$  volume at  $0^\circ$ .

If the law held true for low temperatures at  $-273^\circ$ , the gas would entirely disappear, or occupy no volume at all. This is, of course, impossible ; it has been found that as gases approach the temperature of their liquefying point, they deviate more and more from this law. Nevertheless, for the sake of convenience, the point  $-273^\circ$  is called the *absolute zero*. *Absolute temperature* is, therefore, obtained by adding  $273^\circ$  to the readings of the Centigrade thermometer.

If a volume of gas which measured 273 cc. at  $0^\circ$  Centigrade or  $273^\circ$  *absolute temperature* were heated up  $1^\circ$  to  $274^\circ$  *absolute temperature*, the volume would be 274 cc. ; if it were heated up  $10^\circ$  to an *absolute temperature* of  $283^\circ$ , its volume would be 283 cc. The law of Charles may there-

fore be stated: *The volume of a gas varies directly as its absolute temperature; or*

$$V : v :: T' : t', \text{ whence } V = \frac{vT'}{t'}, \quad (4)$$

where  $T'$  and  $t'$  are the absolute temperatures of the volumes  $V$  and  $v$  respectively.

From Boyle's law we have

$$V : v :: p : P, \text{ whence } V = \frac{vp}{P}, \quad (5)$$

where  $P$  and  $p$  are the pressures under which the volumes  $V$  and  $v$  are measured.

Replacing  $v$  in (4) by the value of  $V$  in (5) for the same reason that  $v$  in (2) was replaced by the value of  $V$  in (1), we get —

$$V = \frac{\frac{vp}{P} T'}{t'} = \frac{vpT'}{Pt'}, \quad (6)$$

a general expression for finding the volume  $V$  of a gas at any absolute temperature  $T'$ , under any pressure  $P$ , when its volume  $v$  is known at any absolute temperature  $t'$ , under any pressure  $p$ .

For the special case of reducing a gas to standard conditions, where  $T' = 273 = 0^\circ \text{C.}$ , and  $P = 760$ , and  $t' = 273 + t$ , formula (6) becomes —

$$V = \frac{vp \cdot 273}{760(273 + t)} = \frac{vp}{760\left(1 + \frac{t}{273}\right)} = \frac{vp}{760(1 + 0.00366 t)},$$

the same as (3).

For example, suppose it were desired to find what volume a gas would occupy when measured at  $50^\circ \text{C.}$  under a pres-

sure of 500 mm. when it was known to occupy 100 cc. at a temperature of  $25^{\circ}$  and a pressure of 750 mm. Here

$$v = 100; p = 750; t' = 273 + 25 = 298;$$
$$P = 500; T' = 273 + 50 = 323;$$

substituting these values in formula (6), we have—

$$V = \frac{100 \times 750 \times 323}{500 \times 298} = 162.58 + \text{cc. } \textit{Ans.}$$

This problem can be solved by using formula (3) instead of (6). Try it.

The following problems will give additional practice in the method of reducing the volumes of gases to standard conditions; all of them may be solved by using either formula (3) or (6), or indeed without any formula at all. The student should not allow himself to substitute blindly numbers for letters in the formula, but should continually bear in mind that these formulæ are merely the laws of Boyle and Charles expressed in algebraic form. A moment's inspection of the problem will always indicate how the proper corrections are to be made without regard to the formula. If, for example, the pressure is increased, we know that the volume must be decreased, and therefore of the two numbers representing the pressure, the larger must be in the denominator. Again, if the temperature is increased, the volume will be increased, and therefore of the two numbers representing the absolute temperature, the larger must be in the numerator; and these considerations enable us at once to state the problem.

#### QUESTIONS IN REVIEW.

1. State clearly the difference between chemical and physical change.
2. Define chemistry.
3. What is meant by measuring a gas under *standard conditions*? Why is it necessary to measure gases under standard conditions?

4. What is the law of Boyle? The law of Charles? What other names are sometimes given to these laws?

5. What is meant by *absolute temperature*? Why is  $-273^{\circ}$  assumed as the absolute zero?

6. A volume of air measures 1 liter at a temperature of  $40^{\circ}$  and under a pressure of 740 mm. What would be its volume under standard conditions?

7. What volume would 10 liters of illuminating gas measured under standard conditions occupy if measured at a pressure of 500 mm. and a temperature of  $100^{\circ}$ ?

8. A volume of gas confined over mercury, the mercury in the tube being 68 mm. above the surface of the mercury in the dish, measured 30 cc.; the barometer was 758 mm.; the temperature was  $25^{\circ}$ . What volume would this gas occupy if measured at a temperature of  $100^{\circ}$  and a pressure of 1000 mm.?

9. When equal volumes of the two gases, oxygen and hydrogen, are weighed under standard conditions, the oxygen is found to be 16 times as heavy as the hydrogen. Assuming the temperature to remain constant, under what pressure would a liter of hydrogen weigh the same as a liter of oxygen weighs under standard conditions? Under what pressure would a liter of oxygen weigh the same as a liter of hydrogen weighs under standard conditions?

10. Assuming the pressure to remain constant, at what temperature would a liter of oxygen have the same weight that a liter of hydrogen has when measured under standard conditions? At what temperature would a liter of hydrogen have the same weight that a liter of oxygen has when measured under standard conditions?

11. At what temperature must oxygen be measured in order that under a pressure of 250 mm. it shall have the same weight, volume for volume, that hydrogen has when measured at  $-40^{\circ}$  under a pressure of 1000 mm.?

## CHAPTER II.

## VAPOR TENSION OF LIQUIDS. — CORRECTIONS FOR VAPOR TENSION IN MEASURING GASES. — QUESTIONS IN REVIEW.

**Vapor Tension of Liquids.** — All volatile liquids evaporate more or less readily at all temperatures. The vapor escaping from their surfaces is capable of exerting a certain pressure, which varies with the nature of the liquid, and increases with the temperature, but in any given liquid is constant at the same temperature.

When this pressure, or tension, just begins to exceed the external pressure upon the liquid, vapor bubbles begin to form in the liquid, and we have the phenomena of *boiling*.

As the vapor pressure of a liquid increases with the temperature, it follows that when the external pressure upon a liquid is reduced, its boiling point is lowered. When we observe the reduced boiling point, and measure the reduced pressure, we know at once the vapor tension of the liquid at the temperature observed, for the vapor tension is equal to the external pressure when the liquid boils. In Exp. 5 we have, then, not only determined the temperatures at which water boils under certain reduced pressures, but have also determined the *vapor tension* of water at those temperatures.

**Corrections for Vapor Tension in Measuring Gases.** — When a gas is measured over water, as is frequently the case, part of the gas volume is made up of water vapor. But if we know the tension of water vapor at the temper-

ature of the experiment, we have a means of knowing how much of the measured volume is made up of water vapor, and how much is made up of the gas we wish to measure.

For example, suppose it is desired to measure the volume of a gas in the closed arm of a U tube over water. (See Fig. 40.)

So long as the water in both arms of the U tube is at the same level, the confined gas is under atmospheric pressure. In the closed arm, however, the gas pressure is made up of the pressure exerted by the water vapor and the pressure exerted by the gas which we wish to measure.



FIG. 40.

Expressing these facts in the form of an equation, we have

$$p'' + p' = p, \text{ whence } p'' = p - p',$$

where  $p''$  is the pressure exerted by the gas to be measured ;  
 $p'$  is the tension of water vapor at the temperature  
 prevailing while the gas is being measured ;  
 $p$  is the barometer reading.

In the formulæ for reducing the volume of a gas to standard or known conditions of pressure and temperature, which have been developed in the previous chapter,  $p$  should be the pressure exerted by the gas to be measured, that is, should be  $p''$ . Hence, when a gas is measured over water, these formulæ become —

$$V = \frac{v(p - p')}{760(1 + 0.00366 t)} \quad \text{and} \quad V = \frac{v(p - p') T'}{P t'}$$

In accurate measurements of gases this correction for the tension of water vapor should always be made when the gases are measured over water, or when measured over mer-

cury, provided the gases are saturated with water vapor. In measuring *dry* gases over mercury, of course this correction is omitted.

A gas should always be measured either dry or else saturated with water vapor; for if it contains water vapor and yet less than is required to saturate it, we cannot tell the amount of water vapor present, and cannot therefore make the proper correction for it. Gases measured over water are always saturated with water vapor. Gases measured over mercury should either be dried before measuring or else saturated with water vapor by introducing a drop or two of water into the vessel under the mercury, and allowing it to rise to the surface of the mercury and remain in contact with the gas for several moments before measuring.

Mercury being a volatile liquid, it follows from what has already been said that when a dry gas is measured over mercury it contains mercury vapor.

The tension of mercury vapor at ordinary temperatures is very slight.

When a gas is measured at atmospheric temperatures, no correction need be made for the tension of mercury vapor, for we measure atmospheric pressure with a mercury barometer, and the barometer stands lower than it should by just the amount of the tension of mercury vapor at the atmospheric temperature.

When gases are measured over mercury at high temperatures, the correction for the tension of mercury vapor should be made, using for  $p'$  in this case the *difference* between the tension of mercury vapor at the temperature of the barometer and the tension of mercury vapor at the temperature at which the gas is measured.

It is also necessary in such cases to take into account the expansion of mercury by heat.



If by any means the temperature of the barometer can be kept the same as that of the measured gas, no corrections for the tension of mercury vapor, or for the expansion of the mercury by heat, will be necessary.

### QUESTIONS IN REVIEW.

1. What is meant by vapor tension? Have all liquids the same vapor tension?

2. Does the vapor tension of water increase by equal amounts for equal increments in the temperature? Compare your results in Exp. 5. See also Table II. in the Appendix.

3. When a liquid boils, what do you know of the tension of its vapor?

4. Why is it necessary to take into account the tension of water vapor when gases are measured over water?

5. Why must a gas be either saturated with water vapor or perfectly dry, in order that its volume may be accurately measured?

6. Why are corrections for the tension of mercury vapor not necessary at ordinary temperatures, but are necessary at higher temperatures? Explain under what conditions these corrections will not be necessary at high temperatures.

7. The volume of a gas measured over water is 15 cc. at  $25^{\circ}$  and 755 mm. pressure. What is its volume under standard conditions? See Table II. in the Appendix for the tension of water vapor at different temperatures.

8. A volume of dry gas measured over mercury at  $50^{\circ}$  and under a pressure of 760 mm. is 100 cc. What would be its volume at the same temperature and pressure if it were saturated with water vapor?

9. A dry gas is measured over mercury at  $250^{\circ}$ , and found to be 200 cc. The barometer at  $20^{\circ}$  reads 740 mm.; reduce the volume to standard conditions, taking into account the tension of mercury vapor.

Tension of mercury vapor at  $20^{\circ}$  is 0.04 mm.

Tension of mercury vapor at  $250^{\circ}$  is 75.75 mm.

## CHAPTER III.

**WATER AS A SOLVENT. — SOLUBILITY OF SOLIDS. — SOLUBILITY OF LIQUIDS. — SOLUBILITY OF GASES. — CRYSTALLIZATION. — WATER OF CRYSTALLIZATION. — QUESTIONS IN REVIEW.**

**Water as a Solvent.** — Water dissolves solids, liquids, and gases, therefore absolutely pure water is very difficult to obtain. Water may be freed from non-volatile solids and liquids by distillation, but unless the process is carried on entirely in platinum vessels, the water will take up impurities from the vessel. In glass vessels, for example, the water dissolves some of the constituents of the glass to a very perceptible degree; and even after it is distilled in platinum, some of the atmospheric gases are dissolved in it: to get rid of these it must be boiled and cooled in a vacuum.

**Solubility of Solids.** — Solid substances differ very much in their solubility in water. There are few substances which, like platinum, are absolutely insoluble; most solids, on the other hand, are more or less soluble in water, varying from substances, like glass, which are only slightly soluble, to substances like calcium chloride and sodium sulphate, large quantities of which can be dissolved. The amount of a substance which will dissolve in a given amount of water varies with the temperature, but is always definite at a given temperature.

In the case of solids, the amount dissolved generally increases with the temperature; there are a few exceptions

to this rule, however. When a solution has dissolved all of a substance that it can dissolve, it is said to be *saturated*. The amount of a substance which is required to saturate 100 parts of water at different temperatures is usually learned by some such method as that employed in Exp.

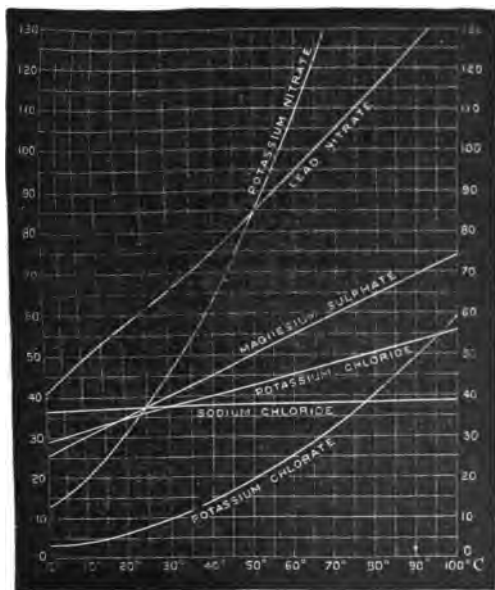


FIG. 41.

14; the larger works on chemistry contain such tables of solubility for many substances.

Instead of tabulating the results obtained by experiment, they can more easily be communicated to the eye by expressing the results in the form of a curve, which at once shows the solubility of the substance in question at any

temperature embraced between the limits of the curve. Figure 41 illustrates a few such curves.

**Solubility of Liquids.** — All liquids are more or less soluble in water. There are first those liquids which, like alcohol, are soluble or miscible with water in all proportions, and second those liquids which, like ether, are only partially or slightly soluble in water. In such cases each liquid is mutually soluble in the other to a greater or less degree. For example, in Exp. 19 the layer of ether on the water contains about 3 per cent of water, and the layer of water below the ether contains about 10 per cent ether. Hence if the total amount of ether put into the water is less than 10 per cent of the amount of water, it will all dissolve, and the two liquids will not separate in layers. In general, increasing the temperature increases the solubility of each liquid in the other, until finally a temperature is reached at which they are soluble or miscible in all proportions; this rule, however, is not without exception, for there are cases where the increase of temperature renders the two liquids less soluble in one another.

Many liquids, such as oils, are only very slightly soluble in water, but it is questionable whether there is a liquid that is absolutely insoluble in water.

**Solubility of Gases.** — All gases are soluble to a greater or less extent in water.

For example : —

- 1 liter of water at 0° dissolves 19.3 cc. of hydrogen;
- 1 liter of water at 0° dissolves 41.14 cc. of oxygen;
- 1 liter of water at 0° dissolves 16.07 cc. of nitrogen;
- 1 liter of water at 0° dissolves 2 liters of chlorine;
- 1 liter of water at 0° dissolves 503 liters of hydrochloric acid;
- 1 liter of water at 0° dissolves 1148 liters of ammonia.

The effect of temperature upon the solubility of gases is

the reverse of that upon the solubility of solids ; the solubility decreases with the increase of temperature.

For example, a liter of water dissolves —

23.34 cc. of air at 10° ;  
19.20 cc. of air at 20° ;  
17.69 cc. of air at 25° ;  
0.00 cc. of air at 100°.

Pressure has only a very slight influence upon the solubility of solids and liquids, but affects very greatly the solubility of gases in water. The temperature remaining constant, the amount of gas which will dissolve in a given amount of water varies directly as the pressure ; that is, twice as much of any gas will be dissolved in a given amount of water under two atmospheres' pressure as will be dissolved under one atmosphere's pressure. This is known as the *Law of Henry*. He stated the law in the following words : " Under equal circumstances of temperature, water takes up in all cases the same volume of condensed gas as of gas under ordinary pressure." This fact enables us to understand the phenomena of effervescent waters, such as soda water, which has dissolved a large amount of carbonic acid gas under pressure ; but when the pressure is removed, much of the gas escapes.

Solutions are being continually formed by the water on the earth's surface, so that all natural waters contain substances in solution which more or less modify their properties, — substances which differ in different localities, so that we have *hard* water and *soft* water, and numberless kinds of mineral water, effervescent springs, etc.

Solution plays an important part in chemical transformations. As seen in Exps. 20, 21, 22, in some cases solution is necessary to chemical action. A solid substance in dis-

solving acquires a mobility of parts, while a gas loses its elasticity, and therefore the conditions are more favorable to a chemical change. Just what occurs when a substance is dissolved is not known.

Indeed, it is only within very recent times that solutions have been studied with that care which their importance warrants, and our knowledge concerning them is now rapidly increasing.

**Crystallization.**—Most substances, when the water in which they are dissolved evaporates, assume definite geometrical forms called *crystals*.

Indeed, "all chemically homogeneous substances when they solidify from a state of vapor, fusion, or solution, tend to assume certain regular polyhedral forms. This tendency is much stronger in some substances than in others, and it varies widely in the same substance under different physical conditions." The form of the crystals is directly dependent upon the nature of the substance composing them, and is therefore valuable for its identification, like any of its physical properties. A thorough knowledge of crystal forms, or *crystallography*, as it is called, is of great use to the chemist.

For this the student is referred to works upon the subject; here only a few points can be mentioned.

When a crystal is formed from an aqueous solution, for example, the smallest crystals, even the smallest particles, possess the essential characteristics of the largest crystals, and the corresponding angles between the corresponding faces are exactly the same.

Crystals not only differ from uncrystallized or *amorphous* bodies in their form, but they also differ in their internal structure. If a homogeneous amorphous body, such as a piece of glass, be examined with regard to its physical properties, such as its power to transmit light, heat, electricity,

its cohesion, its hardness, etc., we will find it equal in all directions. With crystals, however, this is not true. In them we find differences of elasticity, cohesion, hardness, and other physical properties which do not exist in homogeneous substances which are not crystals. It has been found that the distribution of physical properties, like those above enumerated, *is, in crystals, equal along all parallel directions, while, with certain exceptions, it is unequal along directions which are not parallel.* Thus we see that crystals have, besides their form, a certain internal structure which distinguishes them from amorphous bodies.

A piece of glass cut to exactly resemble a crystal would not be a crystal, because it would lack the internal arrangement. Again, a broken fragment of a crystal which was without regular form could be recognized as a crystal by the peculiar variation in its physical properties already referred to.

Although the number of crystal forms which occur in nature and in artificial products is very great, they can all be classified in six groups or systems.

The planes of a crystal are designated by telling where they would cut certain imaginary axes of the crystal.

All crystals which can be referred to three rectangular axes of equal length are said to belong to the *Regular* or *Isometric* system.

Crystals which may be referred to three rectangular axes, two of equal length, and the third either greater or less than these two, are said to belong to the *Tetragonal* or *Quadratic* system.

Crystals that are referred to three rectangular axes, all of unequal length, are called *Orthorhombic* or *Rhombic*.

Crystals referred to three equal axes in one plane, making angles of  $60^\circ$  with one another, and a fourth axis of a dif-

ferent length, vertical to the plane of the other three, are said to be *Hexagonal*.

Crystals which are referred to three axes of unequal length, two of which cut one another obliquely, the third being perpendicular to the plane of the other two, are said to be *Monoclinic*.

And finally, crystals which are referred to three unequal axes, making unequal oblique angles with one another, are said to be *Triclinic*, or to belong to the Triclinic system. All known crystal forms can be referred to one of these six systems; indeed, all crystal forms mathematically possible may be referred to one or other of these systems.

In those systems in which the axes are not of the same length, or in which the angles between the axes are not right angles, the length of the axes and the angles between them vary in crystals of different substances, but are always constant for the same substance.

Substances which chemically are closely related to one another frequently crystallize in the same form; they are then said to be *isomorphous*.

On the other hand, the same substance, in a few cases, crystallizes in forms which belong to different systems, or if in the same system require different axes; such substances are said to be *dimorphous*, or in rare cases, *trimorphous*.

**Water of Crystallization.** — Many solids, when they crystallize from a water solution, take a portion of the water into their crystal structure as *water of crystallization*. This water of crystallization is easily given up again, in some cases even at ordinary temperatures, but in all cases it may be driven off by heat, the crystal form being at the same time lost.

We see, therefore, that for those substances which have water of crystallization the water is essential to the crystal-



line form, but we are ignorant of the part it plays in the structure of the crystal. In crystals of the same kind the amount of water is constant, but in different kinds the amount varies from those crystals which contain no crystal water at all to those crystals which have considerably more than half their weight made up of water. A few substances are di- or tri-morphous, and have different amounts of water in each form.

Chemists have, as yet, no satisfactory explanation why some substances crystallize without water of crystallization, while others need it in order to form crystals. Nor can they, by studying all the other properties of a given compound, predict whether or not it will require water to aid it to form crystals. Or, in other words, the ability to form crystals without the aid of water is, so far as we know, not accompanied by any other properties peculiar to such substances.

#### QUESTIONS IN REVIEW.

1. Why is it difficult to obtain pure water? How can it be done?
2. What is meant by a saturated solution?
3. What can you say in regard to the solubility of liquids in water?
4. What is the effect of temperature upon the solubility of gases in water? What is the effect of pressure?
5. State the law of Henry.
6. How is solution supposed to aid chemical action?
7. What is a crystal? How are crystals formed?
8. How are crystals classified?
9. How could you distinguish between a real crystal and a piece of glass cut in the form of a crystal?
10. What are the six crystalline systems?
11. Define the terms *isomorphous* and *dimorphous*.
12. What is water of crystallization? Do all crystals contain it? Is it constant in amount?

## CHAPTER IV.

ELEMENTS. — MECHANICAL MIXTURES. — CHEMICAL COMPOUNDS.

— INDESTRUCTIBILITY OF MATTER. — DEFINITE PROPORTIONS.

— MULTIPLE PROPORTIONS. — QUESTIONS IN REVIEW.

**Elements.** — When a current of electricity is passed through water (Exp. 24), we find that the water is decomposed or broken up into two gases, hydrogen and oxygen. It has been proven by experiment that the two gases formed weigh together just the same as the water decomposed. When the two gases are reunited, and water formed (Exp. 25), it has been proven that the amount of water formed is exactly equal in weight to the amount of gas which has disappeared, thus proving that these two gases, oxygen and hydrogen, are the only constituents of water.

Now if either oxygen or hydrogen is subjected to the action of electricity, to heat, or to any of the forces that the chemist uses to produce chemical change, it cannot be decomposed. Whenever it undergoes any sort of chemical change, there is an *increase* of weight, thus proving that it has combined with something else, and has not been decomposed into any simpler substance.

Substances like hydrogen and oxygen which chemists are unable to decompose are called chemical *elements*.

It is true that these elementary substances may possibly be compounds of yet simpler substances, but with our present methods we cannot prove this.

Notwithstanding the vast number of substances occurring

in nature which have been examined by chemists, only between sixty and seventy chemical elements have thus far been discovered; all the other substances are compounds or mixtures of two or more of these elements.

Although there are nearly seventy elements known, most of them are of very rare occurrence; by far the larger part of the material which man has been able to examine is made up of not more than twelve different elements, and certainly 99.9 per cent of this material is made up of less than twenty different elements.

Nearly one-half of the solid matter of the earth's crust, so far as can be judged by examination of that part which is accessible, is made up of the one element oxygen, which is the most abundant of all.

The approximate composition of the earth's crust is believed to be about as follows:—

Oxygen . . . . .	44-48%	Calcium . . . . .	1-7%
Silicon . . . . .	22-36%	Magnesium . . . . .	1-3%
Aluminium . . . . .	6-10%	Sodium . . . . .	2-3%
Iron . . . . .	2-10%	Potassium . . . . .	2-3%

Besides forming a large portion of the earth's crust, oxygen forms about one-fifth of the atmosphere and eight-ninths of water. Carbon is the principal element which enters into the structure of living things, while hydrogen, oxygen, nitrogen, phosphorus, and sulphur also are essential constituents of animals and plants.

Nitrogen forms about four-fifths of our atmosphere. On the other hand, many of the rare elements have been found only in small quantities in very rare minerals, which occur in perhaps but one or two known localities; such elements are merely scientific curiosities, their properties have been studied and described by one or two chemists, but the

majority of chemists have never even seen the element or its compounds. Thus it is impossible to state the *exact* number of elements known. There is quite a group of so-called elements which, owing to their great scarcity, have not been thoroughly studied, and it has thus far been impossible to determine whether they are elements or not; for this reason authorities frequently differ when they state the number of elements. A list of sixty-seven elements is given in Table I. of the Appendix. This list includes all those generally recognized as elements.

Many chemical compounds are made up of but two elements, others contain three elements, and some contain four elements; there are comparatively few substances containing more than four elements.

From what has been said, it will be seen that it is by no means necessary to study *all* of the elements to have a good knowledge of chemistry.

**Mechanical Mixture.**—Hydrogen and oxygen may be mixed with one another in every proportion, and so long as the mixture is kept at ordinary temperatures there is no change; it remains merely a mixture of the two gases, the properties of each gas may be recognized in the mixture.

So, too, solid substances in the form of a powder can be intimately mixed in any proportions we choose, yet the properties of each component may be recognized in the mixture; a bit of the powder examined under the microscope would show the little particles of one substance mingled with the little particles of the other. There would be different kinds of particles in it; it would not be homogeneous. For example, suppose fine iron filings and powdered sulphur were thus mixed, the mixture might be so thorough that it would appear to the eye as a homogeneous substance, yet the microscope would show the particles of sulphur and the

particles of iron ; a magnet would immediately attract the iron and leave the sulphur so that the two substances could be separated from one another in that way.

Such substances are called *mechanical mixtures*, because their component parts are not chemically combined with one another, but exist each by itself, and they may be separated one from another by purely mechanical means ; that is, without the aid of chemical change.

**Chemical Compounds.** — If a spark is passed through the hydrogen and oxygen there is an explosion ; a part, or perhaps all, of the gas disappears, and water is formed. Water in its properties in no way resembles the hydrogen and oxygen which have disappeared ; there has been a chemical union of the elements, and a compound is formed. In the mixture the oxygen and hydrogen may be present in any proportions, but in the compound they are present in a single fixed proportion.

We have seen (Exp. 26) that in water there are two volumes of hydrogen for every volume of oxygen, and when these gases unite again they unite in exactly these proportions, no matter how they are mixed ; either gas in excess of these amounts will be left over unchanged after the combination.

So, too, with the iron and sulphur, if they are heated until they are red hot, they will unite chemically with one another ; after the mass is cooled and powdered, it will be found that the powder is neither iron nor sulphur ; it is not magnetic ; it is homogeneous ; that is, the powder when examined with the microscope shows but one kind of particles. Here, too, although the iron and the sulphur may be mixed in all proportions, they combine in one fixed proportion, seven parts, by weight, of iron to four parts of sulphur, and whichever element is present in the mixture before

heating in excess of these proportions will remain over unchanged after the combination.

To summarize, then, a mechanical mixture may have its components present in varying proportions, the properties of the components may be recognized in the mixture; its components can, as a rule, be separated by mechanical means alone. The constituents are not chemically united.

A chemical compound has its component parts present in definite fixed proportions; the properties of the constituents cannot be recognized in the compound; the substance cannot be separated into its constituents without the aid of chemical action. The constituents are chemically united in the compound.

As simple as is the distinction between chemical compounds and mechanical mixtures, and as easy as it is to distinguish between them in most cases, there are substances which require all the knowledge and skill of an experienced chemist to decide to which class they belong.

**Indestructibility of Matter.** — The earlier chemists seem to have tacitly admitted that matter could be destroyed, or could be created, by certain processes; this, no doubt, came about from the fact that in their study of chemical changes they did not investigate the proportions by weight in which the changes took place. Indeed they made little or no use of the balance in their investigations; and, when matter disappeared to their senses, they believed that it was destroyed. When two invisible gases united to form a liquid, or solid, which became then first evident to their senses, they believed that matter had been created. It was the French chemist, Lavoisier (1743-1794), who first introduced the use of the balance into the study of chemical transformations, and who very soon proved the indestructibility of matter. In all chemical changes he found the weight of the material taking

part in the change remained the same ; nothing was gained, nothing was lost.

**Definite Proportions.** — Without any extended investigations upon the subject, Lavoisier seems to have believed in the law of definite proportions ; that is, that *a chemical compound always contains the same constituents in the same proportion by weight*. Berthollet (1748–1822) was the first to throw doubt upon it. He believed that the proportions in which the elements combined were not fixed, but that they varied between certain limits, and that the proportion in any given case depended upon the relative amounts of the elements present, the temperature, and other physical conditions.

Berthollet brought forward experimental evidence to prove his doctrine, and on account of his prominence as a chemist, his views won many supporters.

Proust (1755–1826) opposed the views of Berthollet, and maintained that chemical combination took place in perfectly definite proportions. He proved that many of the substances analyzed by Berthollet were not definite compounds, but really mixtures of two or more compounds.

This discussion lasted through a period of eight years, and most of the chemists of that day were drawn into it. It finally resulted in the complete overthrow of Berthollet's doctrine, and established the law of definite proportions upon a firm basis. All subsequent investigations have confirmed this law.

**Multiple Proportions.** — In many cases the same two elements unite in different proportions to form different compounds. Such is, for example, the case with oxygen and nitrogen, with carbon and oxygen. (See Exps. 93, 98, 108, 116.) An examination of such cases has established the law of multiple proportions, which may be stated as

follows: *When two elements are capable of uniting in different proportions so as to form two or more compounds, the different masses of one element which unite with a fixed mass of the other element bear a simple ratio to one another.*

This law was discovered by the English chemist, John Dalton (1766–1844). Dalton analyzed olefiant gas and marsh gas, both composed of carbon and hydrogen.

Olefiant gas contains 85.7% carbon and 14.3% hydrogen.

Marsh gas contains 75.0% carbon and 25.0% hydrogen.

Upon comparing these numbers he found that the ratio of carbon to hydrogen in olefiant gas was six to one (85.7 : 14.3 :: 6 : 1); whereas in marsh gas it is three to one or six to two (75.0 : 25.0 :: 3 : 1). The mass of hydrogen combined with a given mass of carbon is exactly twice as great in one case as in the other. Struck by this simple relation, Dalton at once examined the oxides of carbon, and found here the same simple ratio.

	Carbon.	Oxygen.
Carbon monoxide contains	42.86 %	57.14 %
Carbon dioxide contains	27.27 %	72.73 %

(42.86 : 57.14 :: 1 : 1.33 and 27.27 : 72.73 :: 1 : 2.66.)

If we represent the amount of carbon in these two compounds by 1, then the amount of oxygen in one case is 1.33 and in the other 2.66; that is, the mass of oxygen combined with a given mass of carbon in carbon dioxide is exactly twice as great as the mass of oxygen combined with the same mass of carbon in carbon monoxide.

Proust, who, before Dalton, had analyzed many such compounds in the course of his discussions with Berthollet, failed to discover this important law, because he expressed his results only in percentage, and did not attempt to com-



pare the masses of one element that combined with a fixed mass of the other in their different combinations.

Although this law was announced by Dalton more than eighty years ago, and since that time many thousands of new compounds have been discovered, not a single exception to the law is known.

### QUESTIONS IN REVIEW.

1. What is a chemical element? How many are there? Are they all of equal importance?
2. What are the most abundant elements?
3. What is a mechanical mixture? Mention some such mixtures that occur in nature.
4. What is a chemical compound? Give some examples.
5. State fully the difference between mixtures and compounds.
6. What is the law of the indestructibility of matter? How is it proven?
7. What is the law of definite proportions? How was it established?
8. What is the law of multiple proportions? Who discovered it?
9. Show how the results obtained in Exps. 93 and 98 illustrate the law of multiple proportions.

## CHAPTER V.

SYMBOLS. — FORMULÆ AND COMBINING WEIGHTS. — EQUATIONS.  
— QUESTIONS IN REVIEW.

**Symbols.** — For the sake of convenience and brevity the chemical elements are frequently designated by abbreviations, or *symbols*. For example, hydrogen is represented by the letter H, oxygen by the letter O. Usually the symbol is merely the first letter of the name of the element. Where there are several names beginning with the same letter, a second letter is added to distinguish them, as for example: B for boron; Ba for barium; Bi for bismuth; Br for bromine; C for carbon; Ca for calcium; Cd for cadmium; Cl for chlorine; Cr for chromium; Co for cobalt; Cs for cæsium; etc.

In a number of cases, however, the symbol is derived from the Latin name instead of the English name, as Au for gold, from the Latin name *aurum*; Fe for iron, from the Latin *ferrum*; K for potassium, from the Latin *kalium*; Na for sodium, from the Latin *natrium*; etc.

A complete list of the chemical elements with their symbols is given in Table I. of the Appendix.

**Formulæ of Compounds; Combining Weights.** — Compounds are expressed by writing the symbols of the elements which compose them side by side. For example, hydrochloric acid being made up of hydrogen, H, and chlorine, Cl, the symbol or formula for hydrochloric acid is HCl. Oxide of mercury being composed of mercury, Hg, and oxygen, O,

its formula is  $\text{HgO}$ . But since these compounds are made up of definite proportions by weight of their component elements, these symbols represent also their weights.

In hydrochloric acid, for example, if we represent the weight of hydrogen present by 1, then the weight of chlorine present is 35.5. Therefore, the symbol H represents one part, by weight, of hydrogen, while the symbol Cl represents a quantity of chlorine 35.5 times as heavy as the amount of hydrogen represented by H.

By the study of different compounds of an element, it has been found possible to select a number such that it, or some simple whole multiple of it, will express the amount, by weight, of that element which is present in any of its compounds. This number, or weight, is called the *combining number*, or *combining weight* of the element.

Since hydrogen has the smallest combining weight of all the elements, its combining weight has been chosen as the unit, and the combining weights of the other elements are given in terms of this unit.

A study of oxygen compounds has led to the adoption of the number 16 as the combining weight of oxygen, likewise 39 has been adopted as the combining weight of potassium, K, and 35.5 for the combining weight of chlorine, Cl. Since potassium chlorate is made up of the elements potassium, chlorine, and oxygen we might at first write its formula  $\text{KClO}$ , as in the case of mercuric oxide and hydrochloric acid, but analysis has shown that potassium chlorate contains 39 parts of potassium, 35.5 parts of chlorine, and 48 parts of oxygen, while the formula  $\text{KClO}$  would indicate 39, 35.5, and 16 parts of potassium, chlorine, and oxygen respectively. To express 48 parts of oxygen we use  $\text{O}_3$ , since O means 16 parts, 3 O will mean 48 parts, and the formula for potassium chlorate becomes  $\text{KClO}_3$ . In the

formulae of those compounds where the amount of the element present is represented by some multiple of its combining weight, the figure expressing this multiple is written after and slightly below the symbol of the element to which it refers ; as  $\text{H}_2\text{O}$  for water,  $\text{H}_2\text{SO}_4$  for sulphuric acid,  $\text{FeCl}_3$  for ferric chloride, and  $\text{Fe}_2(\text{SO}_4)_3$  for ferric sulphate :  $\text{H}_2\text{SO}_4$ , in the case of sulphuric acid, meaning that the amount of hydrogen which is present in the compound is represented by twice its combining weight, and the amount of oxygen present is four times its combining weight.

The methods made use of in deciding upon the combining weights, and the methods of determining the formulae of chemical compounds will be fully explained in Chapters VIII., IX., and X. For the present it will be only necessary for us to understand that the formula of a compound not only tells at a glance of what elements it is composed, but it also tells, provided we know the combining weights of the elements, exactly how much by weight of each element is present in the compound. For example, the formula of mercuric oxide is  $\text{HgO}$ . 200 is the combining weight of mercury, and 16 the combining weight of oxygen ; hence we know that in 216 parts by weight of mercuric oxide there are 200 parts of mercury and 16 parts of oxygen ; or in 100 parts of mercuric oxide there are 92.6 parts of mercury and 7.4 parts of oxygen ; that is, we can calculate its percentage composition from its formula.

Again, taking potassium chlorate,  $\text{KClO}_3$ , and knowing the combining weights of the elements potassium, chlorine, and oxygen to be 39, 35.5, and 16 respectively, we know that  $122.5, (39 + 35.5 + (3 \times 16) = 122.5)$  parts by weight of potassium chlorate are made up of 39 parts potassium, 35.5 parts chlorine, and 48 parts oxygen, or 100 parts by weight of potassium chlorate contain

$$31.8 \left( = \frac{39}{122.5} \times 100 \right) \text{ parts of potassium,}$$

$$29.0 \left( = \frac{35.5}{122.5} \times 100 \right) \text{ parts of chlorine,}$$

and  $39.2 \left( = \frac{48}{122.5} \times 100 \right) \text{ parts of oxygen.}$

Ordinary cane sugar is composed of the three elements carbon, hydrogen, and oxygen, which have the combining weights 12, 1, and 16 respectively. The formula assigned to cane sugar is  $C_{12}H_{22}O_{11}$ .

Hence  $342$  ( $12 \times 12 = 144$ ,  $16 \times 11 = 176$ ;  $144 + 22 + 176 = 342$ ) parts by weight of sugar contain 144 parts of carbon, 22 parts of hydrogen, and 176 parts of oxygen, or 100 parts of sugar contains

$$42.1 \left( = \frac{144}{342} \times 100 \right) \text{ parts of carbon,}$$

$$6.5 \left( = \frac{22}{342} \times 100 \right) \text{ parts of hydrogen,}$$

and  $51.4 \left( = \frac{176}{342} \times 100 \right) \text{ parts of oxygen.}$

By this method it is possible to calculate the percentage composition of any compound from its formula, provided the combining weights of the elements which compose it are known. This is, in reality, reversing the process for determining the formula described in Chapter X.

The determination of the composition of a substance is, in the first instance, of course a matter of experiment; the formula is based upon these experimental results, and may be used to reproduce these results.

**Equations.** — Since it has been found impossible to destroy or create matter by any process whatsoever, it follows that the mass of matter remains the same regardless of the

chemical changes that take place. This fact renders it possible to express chemical changes in the form of an equation. All the substances taking part in a given chemical change are written in the first member of the equation, and all the substances resulting from the change are written in the second member of the equation; each substance being separated from the others by the sign of addition (+), and the two sides of the equation connected by the sign of equality (=). The total amount of each element and the total amount of material must be the same on the two sides of the equation.

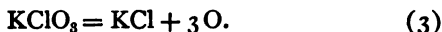
A very simple equation is that representing the decomposition of water by electricity shown in Exp. 26.



The formation of water from its elements, as in Exp. 25, would be expressed by the equation



The formation of oxygen from potassium chlorate, as in Exp. 29, would be expressed



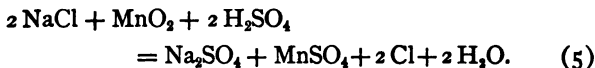
The substance remaining behind after the oxygen is driven off is called potassium chloride, and has the formula KCl.

The formation of hydrochloric acid, HCl, from sodium chloride, NaCl (common salt), and sulphuric acid,  $\text{H}_2\text{SO}_4$  (see Exp. 61), is represented by the equation



Again, the formation of chlorine from sodium chloride,

NaCl, sulphuric acid,  $\text{H}_2\text{SO}_4$ , and manganese dioxide,  $\text{MnO}_2$ , as in Exp. 68, may be expressed by the equation



These equations are short and concise ways of expressing chemical facts :

Equation (1) expresses the fact that water is decomposed, yielding hydrogen and oxygen.

Equation (2) expresses the fact that hydrogen and oxygen unite to form water.

Equation (3) expresses the fact that potassium chlorate is broken up, forming oxygen and potassium chloride.

Equation (4) expresses the fact that sulphuric acid acts upon sodium chloride, and forms hydrochloric acid and sodium sulphate,  $\text{Na}_2\text{SO}_4$ .

Equation (5) expresses the fact that sodium chloride, manganese dioxide, and sulphuric acid react upon one another, producing sodium sulphate, manganese sulphate,  $\text{MnSO}_4$ , chlorine gas, and water.

In each case these facts are expressed in a much briefer way by the equation. But the equation tells us still more than this ; for as we have seen, the formula of a compound shows its component elements, and when the combining weights are known, shows the percentage composition of the compound.

Again, when the equation is *balanced*, — that is, when numerical coefficients are introduced, before the formulæ of some of the compounds, so that the amount of each element shall be the same in both members of the equation, — the equation indicates the exact proportions by weight in which all the substances react with one another in the change.

Let us take for example Equation (5) given above. The combining weights of the elements taking part in this reaction are, sodium, Na, 23; chlorine, Cl, 35.5; manganese, Mn, 55; oxygen, O, 16; hydrogen, H, 1; sulphur, S, 32. Therefore

117 ( $= 2[23 + 35.5]$ ) parts of sodium chloride,  
 87 ( $= 55 + 2[16]$ ) parts of manganese dioxide,  
 and 196 ( $= 2[2 + 32 + 4(16)]$ ) parts of sulphuric acid,  
 reacting upon one another, produce

142 ( $= 2[23] + 32 + 4[16]$ ) parts sodium sulphate,  
 151 ( $= 55 + 32 + 4[16]$ ) parts of manganese sulphate,  
 71 ( $= 2[35.5]$ ) parts of chlorine,  
 and 36 ( $= 2[2 + 16]$ ) parts of water.

( $117 + 87 + 196 = 400$ , and  $142 + 151 + 71 + 36 = 400$ .)

It makes no difference what are the actual amounts that react in any given case, the proportions between them are the same as the proportions between the numbers given above.

If in this case there were 10 grams of sodium chloride, the amounts of manganese dioxide and sulphuric acid necessary to act upon the 10 grams of salt, and the amounts of sodium sulphate, manganese sulphate, chlorine, and water that would be formed, could all be calculated by simple proportions as follows:

117: 87:: 10:  $x$ , where  $x$  is the weight of manganese dioxide;  
 117: 196:: 10:  $x$ , where  $x$  is the weight of sulphuric acid;  
 117: 142:: 10:  $x$ , where  $x$  is the weight of sodium sulphate;  
 117: 151:: 10:  $x$ , where  $x$  is the weight of manganese sulphate;  
 117: 71:: 10:  $x$ , where  $x$  is the weight of the chlorine;  
 117: 36:: 10:  $x$ , where  $x$  is the weight of water.



If instead of 10 grams of salt there were given any amount of any one of the substances in the equation, the corresponding amounts of all the other substances could be calculated by simple proportions, as given above.

Instead of using proportions, these problems may be stated in the form of a fraction, thus: If 117 parts of salt require 87 parts of manganese dioxide, 1 part of salt will require  $\frac{87}{117}$  parts of manganese dioxide; therefore 10 grams of salt will require  $\frac{87}{117} \times 10$  grams of manganese dioxide. In the same manner the other amounts may be determined.

A chemical equation then not only tells us the substances which react upon one another and the substances produced, but it gives the exact composition of all the substances taking part in the reaction, and also shows us the exact proportions by weight in which all the changes take place.

Later we shall see that certain other facts are expressed in the equation also, but for the present we need bear in mind only those here mentioned.

Symbols, formulæ, and equations are of the greatest help to the chemist, and he is continually making use of them. The student should therefore be familiar with their use, and understand just what they are intended to convey to the mind.

On the other hand, it is well to avoid their abuse. In speaking of the elements and compounds, use their names and not their symbols or formulæ. In writing descriptions where it is necessary to refer to elements or compounds, use the names and not the symbols or formulæ, unless it is desired to convey the additional information which is expressed by the formula; then use the name and the formula.

## QUESTIONS IN REVIEW.

1. What are chemical symbols? How are they derived? Of what use are they?

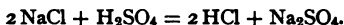
2. What facts can you learn about a compound from its formula?

3. What is meant by the combining weight of an element?

4. What is a chemical equation? What facts are expressed by such an equation?

5. The formula of crystallized copper sulphate is  $\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$ . The combining weights of the elements are: copper, 63; sulphur, 32; oxygen, 16; hydrogen, 1. Calculate the percentage of water of crystallization in copper sulphate. Compare this result with the result obtained in Exp. 9. Calculate the percentage of copper in the crystallized salt. In the anhydrous salt. Calculate the percentage of sulphur in both the crystallized and anhydrous salts.

6. The formation of hydrochloric acid from sodium chloride and sulphuric acid is represented by the equation



The combining weights of the elements in addition to those given in the previous example are: sodium, 23; chlorine, 35.5. What weight and what volume of hydrochloric acid is formed from 25 grams of sodium chloride? How much sulphuric acid would be necessary? How much sodium sulphate would be formed? The specific gravity of hydrochloric acid gas referred to hydrogen is 18.25. One liter of hydrogen weighs 0.0896 gram.

7. The formation of ammonia gas from ammonium chloride (sal ammoniac) and calcium oxide (lime) is represented by the equation—



The combining weight of nitrogen is 14; of calcium is 40. The specific gravity of ammonia gas referred to hydrogen is 8.5. How much lime,  $\text{CaO}$ , and how much ammonium chloride,  $\text{NH}_4\text{Cl}$ , will be required to produce 30 liters of ammonia gas when measured at a temperature of  $27^\circ$  and a pressure of 700 mm.? How much calcium chloride will be formed?

8. Air is 14.4 times as heavy as hydrogen. How much zinc and how much sulphuric acid would be required to produce sufficient hydrogen to just raise a balloon carrying a load of 175 lbs., the balloon itself weighing  $45\frac{1}{2}$  lbs.?

The equation representing the formation of hydrogen from zinc and sulphuric acid is  $\text{Zn} + \text{H}_2\text{SO}_4 = \text{ZnSO}_4 + 2\text{H}$ .

The combining weight of zinc is 65.

## CHAPTER VI.

CHEMICAL NOMENCLATURE. — LAW. — HYPOTHESIS. — THEORY.  
— QUESTIONS IN REVIEW.

**Chemical Nomenclature.** — The names which are given to chemical substances are intended to indicate their composition. The system made use of is not a perfect one, and permits of considerable variation.

Many of the elements have been known for years, and the origin of their names is lost. Those discovered in more recent years have been named by the discoverers according to no prearranged plan.

Some are named from the locality in which they have first been found ; some from a characteristic property, or from the mode of discovery. The nationality of the discoverer has evidently played a part in the naming of the last three elements discovered, — *Gallium*, *Scandium*, *Germanium*.

By common consent names ending in *-ium* have been given to all the recently discovered metals ; for example, sodium, magnesium, and the three recently discovered metals given above.

No general ending is in use for the names of the non-metallic elements, although names of similar elements frequently have a common ending, as the ending *-ine* for the closely allied elements, — fluorine, chlorine, bromine, and iodine ; or the ending *-on* for the group of related elements, — boron, carbon, silicon, etc.

*Compounds.* — When two elements unite to form a com-

pound, the name of the compound is obtained by giving in full the name of the more metallic element followed by the name of the more non-metallic<sup>1</sup> element so modified as to end in *-ide*.

The compound of —

potassium with oxygen is *potassium oxide* ;

sodium with oxygen is *sodium oxide* ;

sodium with sulphur is *sodium sulphide* ;

magnesium with chlorine is *magnesium chloride* ;

calcium with bromine is *calcium bromide*.

In general, the compounds of more basic elements with the following elements are

With	Called	Such as
Hydrogen	Hydrides	Sodium hydride
Oxygen	Oxides	Magnesium oxide
Chlorine	Chlorides	Sodium chloride
Bromine	Bromides	Silver bromide
Iodine	Iodides	Potassium iodide
Fluorine	Fluorides	Calcium fluoride
Sulphur	Sulphides	Lead sulphide
Selenium	Selenides	Potassium selenide
Tellurium	Tellurides	Gold telluride
Phosphorus	Phosphides	Calcium phosphide
Carbon	Carbides	Iron carbide

It frequently happens that an element forms two or more distinct oxides, chlorides, or other compounds ; in such cases a name is given so as to indicate the number of combining weights of oxygen, chlorine, or other elements that

<sup>1</sup> The terms *metal*, *non-metal*, *acid*, *base*, and *salt*, which are frequently used in this chapter, are fully explained in Chapter XIII.

are supposed to be present, as carbon *monoxide*,  $\text{CO}$ , carbon *dioxide*  $\text{CO}_2$ ; sulphur *dioxide*,  $\text{SO}_2$ , sulphur *trioxide*  $\text{SO}_3$ ; phosphorus *trichloride*,  $\text{PCl}_3$ , phosphorus *pentachloride*,  $\text{PCl}_5$ . The lower oxides are also sometimes called *suboxides*, and the higher oxides *peroxides*.

A few metals combine with two different amounts of oxygen, chlorine, and other elements, to form two distinct series of compounds; that series containing the *least* oxygen, chlorine, or other elements is designated by the ending *-ous* given to the name of the metal; that series containing the *most* oxygen, chlorine, or other elements is designated by the ending *-ic*.

Thus:—

mercurous chloride, $\text{HgCl}$ ;	mercuric chloride, $\text{HgCl}_2$ ;
mercurous oxide, $\text{Hg}_2\text{O}$ ;	mercuric oxide, $\text{HgO}$ ;
antimonous chloride, $\text{SbCl}_3$ ;	antimonic chloride, $\text{SbCl}_5$ .

In a few cases the endings *-ous* and *-ic* are used with the Latin name.

The two chlorides of iron, for example, are

ferrous chloride, $\text{FeCl}_2$ ;	ferric chloride, $\text{FeCl}_3$ .
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The two chlorides of tin are

stannous chloride, $\text{SnCl}_2$ ;	stannic chloride, $\text{SnCl}_4$ .
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*Acids*.—All acids contain hydrogen, and most of them contain oxygen, united with some other element. Their names are derived usually from the characteristic element of the acid with which the oxygen and hydrogen are united, and usually end in *-ic*; thus:—

$\text{H}_2\text{SO}_4$	... .. sulphuric acid;
$\text{HNO}_3$	... .. nitric acid;

$\text{HClO}_3$	chloric acid ;
$\text{H}_3\text{PO}_4$	phosphoric acid ;
$\text{H}_2\text{CO}_3$	carbonic acid.

In a number of cases the same elements form two acids ; then the one containing the most oxygen has the name ending in *-ic*, and the one containing less oxygen has the name ending in *-ous* ; as —

$\text{H}_2\text{SO}_3$	sulphurous acid ;
$\text{HNO}_2$	nitrous acid ;
$\text{HClO}_2$	chlorous acid ;
$\text{H}_3\text{PO}_3$	phosphorous acid.

In case there is a third acid containing still less oxygen than the one having a name ending in *-ous*, the prefix *hypo-* (from *ὑπό*, *under*, *beneath*) is added : —

$\text{H}_2\text{SO}_2$	<i>hyposulphurous</i> acid ;
$\text{HNO}$	<i>hyponitrous</i> acid ;
$\text{HClO}$	<i>hypochlorous</i> acid.

In case there is an acid containing still more oxygen than the acid having the name ending in *-ic*, the prefix *per-* (from the Latin *per*, signifying completeness) is added ; as —

$\text{HClO}_4$	<i>perchloric</i> acid.
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Those acids which contain no oxygen have names ending in *-ic*, and have the prefix *hydro-* added to their names ; as —

$\text{HCl}$	<i>hydrochloric</i> acid ;
$\text{HBr}$	<i>hydrobromic</i> acid ;
$\text{HCN}$	<i>hydrocyanic</i> acid ;
$\text{H}_2\text{S}$	<i>hydrosulphuric</i> acid.

*Bases.*—Bases or compounds of the metals with oxygen and hydrogen are known as *hydroxides* (also sometimes as *hydrates*) :—

NaOH . . . . .	sodium hydroxide ;
KOH . . . . .	potassium hydroxide ;
Ca(OH) <sub>2</sub> . . . .	calcium hydroxide ;
Fe(OH) <sub>2</sub> . . . .	ferrous hydroxide ;
Fe(OH) <sub>3</sub> . . . .	ferric hydroxide.

*Salts.*—When an acid and a base act upon one another, a *salt* is produced ; or when an acid acts upon a metal, the hydrogen of the acid is replaced by the metal, and a *salt* is formed. Salts derived from acids having names ending in *-ous*, are given names ending in *-ite* ; as—

sulphurous acid	gives sulphites ;
hyposulphurous acid	gives hyposulphites ;
nitrous acid	gives nitrites ;
hyponitrous acid	gives hyponitrites ;
chlorous acid	gives chlorites ;
hypochlorous acid	gives hypochlorites ;
phosphorous acid	gives phosphites.

If the salt is derived from an acid whose name ends in *-ic*, the name of the salt ends in *-ate*. For example :—

sulphuric acid	gives sulphates ;
nitric acid	gives nitrates ;
phosphoric acid	gives phosphates ;
chloric acid	gives chlorates ;
perchloric acid	gives perchlorates ;
acetic acid	gives acetates ;
carbonic acid	gives carbonates ;
chromic acid	gives chromates.

The naming of the salts derived from the *hydro* acids forms an exception to this rule. The salts of hydrochloric acid are not called *hydrochlorates*, because the same salt is formed by the direct union of the metal with chlorine, and is therefore, as we have seen, called a *chloride*; the name *hydrochlorate*, being superfluous, is not used.

For the same reasons

hydrobromic acid	gives bromides ;
hydriodic acid	gives iodides ;
hydrocyanic acid	gives cyanides ;
hydrosulphuric acid	gives sulphides.

The name of any particular salt is formed by giving the name of the metal which it contains, and following this with the name of the class to which it belongs.

As the following examples will illustrate :—

$\text{KNO}_3$	. . . . .	potassium nitrate ;
$\text{KNO}_2$	. . . . .	potassium nitrite ;
$\text{Na}_2\text{SO}_4$	. . . . .	sodium sulphate ;
$\text{Na}_2\text{SO}_3$	. . . . .	sodium sulphite ;
$\text{KClO}_3$	. . . . .	potassium chlorate ;
$\text{KClO}_4$	. . . . .	potassium perchlorate ;
$\text{CaCl}_2$	. . . . .	calcium chloride ;
$\text{FeSO}_4$	. . . . .	ferrous sulphate ;
$\text{Fe}_2(\text{SO}_4)_3$	. . . . .	ferric sulphate ;
$\text{Ba}(\text{NO}_3)_2$	. . . . .	barium nitrate ;
$\text{NaClO}$	. . . . .	sodium hypochlorite.

With a little experience and by the aid of the principles of nomenclature here given, the student will be able to interpret most chemical names met with in inorganic chemistry. Many substances have other names besides their chemical names, which are frequently used in commerce,



by pharmacists, by physicians, and by chemists themselves. Such are, for example : —

Glauber's salt	for sodium sulphate, $\text{Na}_2\text{SO}_4 \cdot 10 \text{H}_2\text{O}$ ;
Epsom salt	for magnesium sulphate, $\text{MgSO}_4 \cdot 7 \text{H}_2\text{O}$ ;
calomel	for mercurous chloride, $\text{HgCl}$ ;
corrosive sublimate	for mercuric chloride, $\text{HgCl}_2$ ;
lunar caustic	for silver nitrate, $\text{AgNO}_3$ ;
caustic soda	for sodium hydroxide, $\text{NaOH}$ ;
caustic potash	for potassium hydroxide, $\text{KOH}$ ;
aqua fortis	for nitric acid, $\text{HNO}_3$ ;
muriatic acid	for hydrochloric acid, $\text{HCl}$ ;
oil of vitriol	for sulphuric acid, $\text{H}_2\text{SO}_4$ ;
blue vitriol	for copper sulphate, $\text{CuSO}_4 \cdot 5 \text{H}_2\text{O}$ ;
and many others.	

### **Law ; Hypothesis ; Theory.**

*Law.*—A natural law is the regular sequence in which certain results follow certain causes.

Natural laws are learned by careful observation of causes and results.

Our announcement of natural laws is frequently at fault through imperfect observation or through insufficient observation.

Natural laws are usually announced after a comparatively few observations, and farther observation may more fully confirm such announcements or may prove them false.

Dalton announced the law of multiple proportions after he had examined a comparatively small number of compounds ; that it applied to similar cases that had not been examined he could not of course know, but he believed that it did. Subsequent observation has confirmed this law in the most complete manner, so that now chemists have come almost to believe that no exception to the law could exist.

Few are so fortunate in their announcement of laws as was Dalton in this case. It much more frequently occurs that the first statement of the law is much modified by subsequent observations.

To properly estimate the value of the statement of any law it is necessary to know something about the number and nature of the observations upon which it is based.

*Hypothesis.*—It is one thing to know a law and quite another to know the *cause* of the law.

For example, we may be perfectly familiar with the laws of definite and multiple proportions without being able to answer the question: Why do the elements combine with one another according to the laws of definite and multiple proportions? Whenever a natural law is discovered it at once suggests the question: *Why is this so?* To answer this question effort is made to imagine a condition of things in nature which, if it did exist, would cause the law to follow as a necessity. Such a guess at the condition of things is called an *hypothesis*.

Any hypothesis put forth to explain a law will indicate that, if it is true, certain other phenomena must follow as well as those which it was put forward to explain.

If upon examination it is found that these other phenomena do not follow, the hypothesis is false and is rejected.

But on the other hand if these phenomena do follow, they furnish additional evidence in favor of the correctness of the hypothesis.

Thus an hypothesis may be tested, and when found to be in accordance with a great number of facts it is called a *theory*,<sup>1</sup> and is used in the farther development of the science.

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<sup>1</sup> The terms *hypothesis* and *theory* are frequently used as synonyms, but the tendency of present scientific writers is to use them in the sense here given, though the boundary line between hypothesis and theory is necessarily indistinct.

As soon as a fact is discovered which could not exist if the theory were true, the theory must be modified to accommodate this new fact ; or, if this is impossible, the theory must be given up. In any science, therefore, in which discoveries are being made, the theories are being continually modified, some being rejected, others extended and made still closer approximations to the truth.

Many scientific theories are far from being satisfactory ; they are retained only because, as yet, with the facts at hand, no one has been able to suggest better ones. Other theories are so satisfactory, and are supported by such a large array of facts, that they are looked upon almost as laws. A theory can never be actually proven to be true, but the evidence in its favor may accumulate to such an extent that to doubt it is absurd.

To estimate the probability of the correctness of a theory, we must be familiar with all the evidence upon which it is based.

*Use of Theories.* — Theories in a science continually suggest new lines of investigation. The efforts to overthrow or establish them lead to a vast amount of experimental work in which many new facts are learned, and the science thereby advanced. Even false theories may be of positive value to a science by the stimulus they give to research.

*Harm of Theories.* — The harm a theory may do a science comes from the tendency it has to cause scientists to interpret all new facts according to some theory, when another interpretation might perhaps be more satisfactory. Theories do harm to a science in the minds of those who are ignorant of the true relation of theories to a science ; such persons feel that when a scientific theory is disproved, a blow has been struck at the science itself, when in reality it has been advanced and strengthened by the change.

The student of chemistry should have clearly in mind the facts that are to be interpreted by a chemical theory. He should examine the evidence for himself.

He should remember that a theory is, after all, only the most probable explanation based upon the evidence at hand, and that new evidence may at any time be obtained which renders the old explanation improbable.

### QUESTIONS IN REVIEW.

1. Explain the significance of the endings *-ous* and *-ic* in naming different compounds of the same metal.
2. Explain the significance of the endings *-ic* and *-ous* when applied to the names of acids.
3. Explain the meaning of the endings *-ite* and *-ate* when applied to the names of classes of salts.
4. Explain the use of the prefixes *hypo-*, *per-*, and *hydro-*.
5. Give names to the compounds represented by the following formulæ: —

KClO,	NaNO <sub>2</sub> ,	Na <sub>2</sub> SO <sub>3</sub> ,	HgNO <sub>3</sub> ,
KClO <sub>2</sub> ,	KNO <sub>3</sub> ,	K <sub>2</sub> SO <sub>4</sub> ,	Hg(NO <sub>3</sub> ) <sub>2</sub> ,
KClO <sub>3</sub> ,	CaSO <sub>3</sub> ,	FeCl <sub>2</sub> ,	CuCl,
NaClO <sub>4</sub> ,	CaSO <sub>4</sub> ,	FeCl <sub>3</sub> ,	CuCl <sub>2</sub> .

6. Give the distinction between the terms *law*, *hypothesis*, and *theory*.
7. What are the uses of a theory to a science?

## CHAPTER VII.

THE ATOMIC THEORY. — ATOM. — MOLECULE. — QUESTIONS  
IN REVIEW.

**The Atomic Theory.** — Dalton was not content with the discovery of the law of multiple proportions, but sought the cause of the law. He sought an answer to the question: Why do the elements combine according to the laws of definite and multiple proportions? To answer this, he put forward his now celebrated atomic hypothesis. In this hypothesis Dalton assumes that all chemical elements are made up of minute homogeneous particles, which have a definite and a constant weight for each element, and that in the course of a chemical change these particles, which he called *atoms* (from *ἄτομος*, uncut or indivisible), remain undivided.

Farther he assumed that chemical compounds are formed by the union of the atoms of the different elements in the simplest numerical proportions; that is, one atom of one element uniting with one of another element, or one atom of one element uniting with two of another, or one with three, or one with four, or two with three, or two with five, etc. This hypothesis accounts in a very satisfactory manner for the laws of definite and multiple proportion. Take, for example, hydrochloric acid; suppose the hydrogen and chlorine combine atom with atom, and the chlorine atom is 35.5 times as heavy as the hydrogen atom, then hydrochloric acid *must* have the composition 1 part by weight of

hydrogen to 35.5 parts by weight of chlorine. Take the two oxides of carbon examined by Dalton; suppose the weight of the atoms of carbon and oxygen are to each other as 1 to 1.33, and that in carbon monoxide we have one atom of carbon and one atom of oxygen, then the composition of carbon monoxide *must* be 1 part by weight carbon to 1.33 parts oxygen. Suppose that in carbon dioxide there are one atom of carbon and two atoms of oxygen; then the composition of carbon dioxide must always be 1 part, by weight, carbon to 2.66 parts oxygen. That is, the facts which we observe to be true *must* follow as a necessity if this hypothesis is true. This hypothesis was found to be in accordance with all the facts known at the time, and was very generally accepted by chemists, as it gave such a satisfactory explanation of the phenomena observed. The atomic theory harmonizes with all the discoveries that have been made since it was first introduced. There is, perhaps, no other scientific theory that is supported by such an array of facts. Indeed, it seems impossible that a false theory could so well account for the phenomena which are observed. The atomic theory forms the basis of all chemical speculation, yet it is well to bear in mind that if some future discovery should prove the atomic theory to be false, it would not change a single chemical fact. It might, it is true, completely revolutionize our *conception* of chemical facts.

The atomic theory introduced by Dalton was no new idea; the ancient philosophers and schoolmen had before discussed the constitution of matter, some maintaining that matter was infinitely divisible, while others claimed that there was a limit beyond which division could not go. Each side defended its belief with keen arguments and great logical acumen; but, of course, so long as matter itself was not studied, no conclusions could be reached.

Whether or not Dalton was familiar with these discussions of the ancients, it is impossible to say; but certainly to Dalton belongs the credit of having given a definite meaning to the atomic theory, since he presented the evidence in its favor so clearly that it was at once quite generally accepted.

As the atoms have, according to this theory, a definite and constant weight, Dalton at once set himself the task of determining their weights.

He saw that to determine their absolute weights would be impossible; but he thought their relative weights, as compared with one another, might be determined by simply analyzing the compounds they formed with one another. He adopted as his unit the weight of an atom of hydrogen, that being the lightest of all. When two elements formed but one compound, Dalton believed that one atom of one element united with one atom of the other element. This was, however, an arbitrary assumption on his part, since there was no evidence for it whatever, — an assumption that we now believe to be false. Nevertheless, with this assumption we can see how Dalton proposed to determine the atomic weights. Water, for example, being at that time the only compound of hydrogen and oxygen known, he assumed that it contained one atom of hydrogen and one atom of oxygen; and since by analysis water was known to have eight times as much oxygen by weight as it has hydrogen, it follows that if the atomic weight of hydrogen is 1, the atomic weight of oxygen must be 8.

Carbonic oxide, Dalton assumed, contained one atom of carbon and one atom of oxygen, and since in this gas 6 parts of carbon are united with 8 parts of oxygen, 8 being the atomic weight of oxygen, 6 is the atomic weight of carbon. In carbonic acid gas there are 6 parts of carbon, by weight, to 16 parts of oxygen; in this compound there must be, therefore, one atom of carbon and two atoms of oxygen.

Here, again, was an arbitrary assumption on the part of Dalton ; although we now believe he was right in this assumption, he would at that time, as far as the evidence was concerned, have been fully as much justified in assuming that carbonic acid gas was the simple compound containing one atom of carbon and one atom of oxygen, and since in this compound 3 parts of carbon are united with 8 parts of oxygen, 3 would have been the atomic weight of carbon, and carbonic oxide gas, containing 6 parts of carbon united with 8 parts of oxygen, would then be made up of two atoms of carbon and one atom of oxygen.

The question whether the atomic weight of carbon was to be 6 or 3 really depended upon an arbitrary assumption.

This arbitrary nature of the atomic weights determined by Dalton apparently in no way disturbed him. Other chemists, however, saw the difficulties, but could see no way of avoiding them ; that is, they could see no way of determining the real atomic weights, because they could not determine how many atoms of the different elements were present in a compound. Some chemists, therefore, gave up the whole theory. Wollaston (1766-1828) proposed that, instead of attempting to determine the atomic weights, chemists should determine the equivalent weights of the elements, and defined equivalent weights of the elements as the weights of the elements that combined with one another, thus abandoning the atomic weights altogether. But how to decide which was the true equivalent weight in cases like carbonic oxide and carbonic acid gas, where each element may have two different equivalent weights, he did not tell, and in his own experimental work he altogether avoided such compounds.

Wollaston's equivalent weights were, in fact, identical with Dalton's atomic weights. Instead of helping to clear up the



confusion he only added to it ; for he introduced the idea of equivalent weight, which is in reality different from the atomic weight, and yet gave to it a meaning precisely the same as Dalton had given to the atomic weight.

To follow out thoroughly the development of the methods of determining the atomic weights, while a very interesting and profitable study, does not properly fall within the limits of the present work.

Here it is desired only to give a clear idea of the problem that is presented in the determination of the atomic weights, to point out some of the difficulties, and finally to show how these difficulties have been overcome. That the problem was a difficult one is made evident by the length of time that it was under discussion, and by the fact that it attracted and held the attention of men who for clearness of conceptions and keenness of intellect have had few superiors in the nineteenth century.

The atomic theory was announced to the world by Dalton in 1808 ; Dalton's friend, Thompson, had, with his permission, published an outline of the theory a year or so before this. The discussion over the atomic weights began at once, and lasted until about 1860. During this time the effort to get at the true atomic weights was almost continuous. But, owing to lack of agreement upon the methods to be employed, many elements had two or three atomic weights, which were used by different writers, and, in consequence,<sup>1</sup> different formulæ were used for the same compounds, all of which led to much confusion.

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<sup>1</sup> The reader will, perhaps, before this, have been convinced that the combining weights spoken of in Chapter V., in the light of the atomic theory, are really the atomic weights, and that the symbol of the element may therefore represent an atom of the element.

At present, however, the methods are so well based upon experimental facts that they all lead to the same results.

The only problem now presented by the atomic weights is, that of obtaining, by more careful experiment, more exact numbers than we now have. When we say that the atomic weight of oxygen is 16 it is not strictly true. The most careful experiments prove that it is 15.96, or perhaps even less. For ordinary purposes the nearest whole number is usually taken, but in accurate calculations the most exact number obtainable should be used.

In Table I. of the Appendix are given the numbers deduced from the present exact experimental data. As the methods of analysis are improved, these numbers may be still more accurately determined; but the changes in their value in most cases will be very slight.

The settling of the question of atomic weights no doubt accounts in a large measure for the wonderful progress made in chemistry within the last thirty-five years. So long as one of the fundamental problems of the science was unsolved, it attracted much attention and time; but so soon as that was settled, chemists at once turned their attention to other and new problems.

**Molecules.** — When an atom of hydrogen and an atom of chlorine unite, they form the smallest part of hydrochloric acid that can be conceived; yet it is not an atom, for by definition an atom is indivisible, and this smallest part of hydrochloric acid can be divided into an atom of chlorine and an atom of hydrogen; the term *molecule* is used to express this smallest amount of hydrochloric acid. The weight of a molecule or the *molecular weight*, as it is called, is, of course, equal to the weight of all the atoms that compose it.

Although chemists and physicists are continually speaking

of atoms and molecules, — of their weight, and even of their size, — and although many elementary text-books speak of them as though their existence was as evident to the senses as the existence of the sun and other heavenly bodies, the student should remember that atoms and molecules are merely conceptions of the mind used to explain certain natural phenomena. No one has ever seen an atom or a molecule ; and if recent deductions are correct, which apparently limit the possibilities of the microscope, they never will be seen.

Neither atoms nor molecules can have any meaning whatever unless we are acquainted with some of the reasons for the assumption of their existence.

To know that an atom is "a portion of matter that cannot be divided," or that a molecule "is the smallest amount of a substance that can exist by itself," without knowing anything else about them, can be of no possible value nor can it convey any clear idea to the mind.

#### QUESTIONS IN REVIEW.

1. What is the atomic theory? What facts was it intended to explain?
2. Show clearly how the theory explains the facts that it was intended to explain.
3. How did Dalton propose to determine the atomic weights.
4. Point out clearly the difficulties that made other chemists unwilling to accept Dalton's atomic weights.
5. Explain the meaning of the terms *molecule* and *molecular weight*.

## CHAPTER VIII.

AVOGADRO'S THEORY. — DETERMINATION OF MOLECULAR WEIGHTS WITH THE AID OF AVOGADRO'S THEORY. — DETERMINATION OF ATOMIC WEIGHTS WHEN MOLECULAR WEIGHTS ARE KNOWN. — RAOULT'S LAW. — DETERMINATION OF MOLECULAR WEIGHT WITH THE AID OF RAOULT'S LAW. — QUESTIONS IN REVIEW.

**Avogadro's Theory.** — Since all gases behave alike toward changes of temperature and pressure, it is believed that *under the same conditions of temperature and pressure equal volumes of all gases contain the same number of molecules.* This theory was first advanced in 1811 by the Italian physicist Avogadro, and hence is known as Avogadro's Theory. In France, a few years later, Ampère put forward the same theory, and it is also sometimes referred to as Ampère's Theory. In 1811, when the theory was announced, it was not adopted by chemists, because at that time the chemical evidence that bore upon this point was of a very conflicting character, as is pointed out in Chapter XI. Indeed, the theory was not generally admitted by chemists until much later, but it has been found to be in accordance with all facts now known. As soon as it was seen to be satisfactory, it became a help in the determination of the atomic weights.

If this theory is true, it follows that the ratio between the specific gravities of two gases is the same as the ratio between their molecular weights. For the ratio between

the specific gravities is the ratio between the weights of equal volumes<sup>1</sup> of the gases. Since in each volume there are the same number of molecules,  $n$  molecules for example, the ratio between the specific gravities is the ratio between the weights of  $n$  molecules of each gas, which must be the same as the ratios between the weights of one molecule of each gas. Expressed in the form of an equation, this is

$$\frac{\text{sp. gr.}}{\text{Sp. Gr.}} = \frac{n \cdot m}{n \cdot M} = \frac{m}{M},$$

$m$  and  $M$  being molecular weights, capital letters referring to one gas and small letters to another.

The specific gravity of gases can be determined by experiment; hence if we knew the molecular weight of *one* gas, it would be possible to calculate the others from the simple proportions given above, but unless we do know the molecular weight of *one* gas this will be impossible.

We must, then, obtain the molecular weight of some gas by other methods, and that will enable us to reach conclusions concerning the molecular weights of other gases.

If the formula of hydrochloric acid is  $\text{HCl}$ , its molecular weight must be 36.5, since, calling the amount of hydrogen present 1, the amount of chlorine present is represented by 35.5. The reason for believing that hydrochloric acid contains one atom of hydrogen and one atom of chlorine, is that it is the *simplest* assumption, and is in accordance with all the properties of hydrochloric acid.

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<sup>1</sup> When volumes of different gases are compared, unless otherwise stated, it is to be understood that such comparisons are always made with the assumption that the gases are measured under the same conditions of temperature and pressure. When the weight of a definite volume of a gas is given, this weight refers to the volume of the gas under *standard conditions*.

If it had more than one atom of hydrogen it might yield, by certain chemical changes, salts in which only a part of the hydrogen has been replaced. Sulphuric acid, phosphoric acid, and, in fact, all compounds which we believe to contain more than one atom of hydrogen, do, by certain chemical processes, yield derivatives which, while they contain some hydrogen, contain less than was in the original compound. Now in hydrochloric acid this is impossible; if any of the hydrogen is driven out of hydrochloric acid, it is all driven out. In other words, the amount of hydrogen in hydrochloric acid cannot be subdivided by any known process; it is therefore, according to the definition, one atom.

For similar reasons, since the amount of chlorine in hydrochloric acid cannot be subdivided, it is believed that it contains but one atom of chlorine. The molecular weight of hydrochloric acid must therefore be 36.5.

**Determination of Molecular Weights with the Aid of Avogadro's Theory.**—If it is desired to determine the molecular weight of any gas, it is possible to find it from the simple proportion

Sp. gr. of gas : Sp. gr. HCl :: M. wt. of gas : M. wt. of HCl ;

or, 
$$\text{M. wt. of gas} = \frac{\text{Sp. gr. gas} \times \text{M. wt. of HCl}}{\text{Sp. gr. HCl}}$$

Let us first assume that the unit of specific gravity used is hydrogen.<sup>1</sup> Then as hydrochloric acid is 18.25 times as heavy as hydrogen, our formula above given is —

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<sup>1</sup> There are two units commonly used in the determination of the specific gravity of gases; one is hydrogen, and the other is air. Of these units, hydrogen is to be preferred on account of the simplicity of the relation which exists between the molecular weight of any gas, and its specific gravity referred to hydrogen.

$$\text{M. wt. of gas} = \frac{\text{Sp. gr. of gas} \times 36.5}{18.25} = \text{Sp. gr. gas} \times 2.$$

Or, in general, we have the simple relation that the molecular weight of a gas is twice its specific gravity referred to hydrogen.

The specific gravity of hydrochloric acid referred to air is 1.267; the formula given above then becomes —

$$\text{M. wt. of gas} = \frac{\text{Sp. gr. of gas} \times 36.5}{1.263} = \text{Sp. gr. of gas} \times 28.88.$$

Or, in general, we have the relation that the molecular weight of a gas is equal to its specific gravity referred to air multiplied by 28.88. In other words, 28.88 is the ratio of the molecular weight of any gas to its specific gravity referred to air.

Again air is 14.44 times as heavy as hydrogen; consequently if we multiply the specific gravity of a gas referred to air by 14.44, it gives us the specific gravity of the gas referred to hydrogen; but we have seen that the molecular weight of a gas is twice its specific gravity referred to hydrogen.

Or,  $\text{Sp. gr. r. air} \times 14.44 \times 2 = \text{M. wt.};$

hence,  $\text{Sp. gr. r. air} \times 28.88 = \text{M. wt.},$

which is the conclusion reached above.

The theory of Avogadro enables us to determine the molecular weights of compounds which are gases, or which can be converted into gases without decomposition, by simply multiplying their specific gravities referred to hydrogen by 2 or their specific gravities referred to air by 28.88.

**Determination of Atomic Weights when Molecular Weights are Known.**— But the molecular weights do not give us directly the atomic weights ; for, plainly, unless we know the number of atoms in the molecule, we must still be in doubt as to the atomic weights.

While it is impossible, in this way, to get direct evidence concerning the number of atoms in a molecule, it is possible to get indirect aid in the determination of the atomic weights.

All the gaseous or volatile compounds of any given element may be analyzed, and their percentage composition determined. Their molecular weights may also be determined by the method just given. Knowing the percentage composition and the molecular weight of the different compounds, the weight of the element in question present in each molecule of its different compounds can be determined.

The weight of the smallest amount of the element present in any molecule of its different compounds is taken as its atomic weight ; for, according to our definition, an atom is the smallest amount of an element which enters into chemical compounds.

Suppose it were desired to find the atomic weight of oxygen from the data furnished by the following compounds :—



Substance.	Per Cent of Oxygen.	Per Cent of Other Elements.	Molecular Weight. <sup>1</sup>	Weight of Oxygen in Molecule.	Weight of Other Elements in Molecule.
Water . . . . .	88.88	11.12	18	16	2
Carbon monoxide . . .	57.14	42.86	28	16	12
Carbon dioxide . . . .	72.73	27.27	44	32	12
Nitrous oxide . . . . .	36.36	63.64	44	16	28
Nitric oxide . . . . .	53.33	46.67	30	16	14
Sulphur dioxide . . . .	50.00	50.00	64	32	32
Sulphur trioxide . . . .	60.00	40.00	80	48	32
Methyl alcohol . . . . .	50.00	{ 37.50 12.50	32	16	{ 12 4
Ordinary alcohol . . . .	34.78	{ 52.18 13.04	46	16	{ 24 6
Ether . . . . .	21.62	{ 64.87 13.51	74	16	{ 48 10

Upon examination of that column of figures which expresses the weight of oxygen in the molecule, we find that

<sup>1</sup> The specific gravity of steam, when determined with the greatest care, rarely gives exactly 18 as the molecular weight of steam; but for the present purpose this is of no consequence. We know from analysis that water contains 1 part of hydrogen to 8 parts of oxygen. If water contains one atom of hydrogen, the molecular weight of water must be  $1 + 8 = 9$ . If it contains two atoms of hydrogen, then the proportion instead of being 1 to 8 is 2 to 16, and the molecular weight of water is 18. If it contains three atoms of hydrogen, the proportion would be 3 to 24, and the molecular weight of water would be 27, and so on. We know by analysis, then, that the molecular weight of water is 9, or some multiple of 9. Now if the specific gravity of steam gives us a number somewhere near 18 as the molecular weight of water, we know that the true number is 18, and not 9 or 27.

The same is true of the other compounds in the table. The molecular weights given above are not the actual numbers obtained by multiplying the specific gravity of the gas by 2, or by 28.88, as the case may be, but are these numbers corrected by the results of analysis, because our methods of analysis are capable of giving more exact figures than are given by our methods for the determination of the specific gravity of gases.

the smallest weight in any molecule is 16, and that the other weights are all some whole multiple of 16. If upon the examination of all the volatile compounds of oxygen in this way, it is found that none of these contain less than 16 parts of oxygen, we are justified in assuming that 16 is the atomic weight of oxygen. If, however, there should be discovered a compound which, when examined by the methods outlined, should be found to contain 8 parts of oxygen; 8 would thereafter be assumed as the atomic weight of oxygen. But so long as such a compound is unknown, 16 is assumed to be the atomic weight of oxygen.

In a similar manner, the atomic weights of other elements which have volatile or gaseous compounds may be determined. The following compounds of carbon give data for that element:—

Substance.	Per Cent of Carbon.	Per Cent of Other Elements.	Molecular Weight.	Weight of Carbon in Molecule.	Weight of Other Elements in Molecule.
Carbon monoxide . . .	42.86	57.14	28	12	16
Carbon dioxide . . . .	27.27	72.73	44	12	32
Olefiant gas . . . . .	85.72	14.28	28	24	4
Marsh gas . . . . .	75.00	25.00	16	12	4
Cyanogen . . . . .	46.15	53.85	52	24	28
Hydrocyanic acid . . .	44.44	{ 51.85 3.71	27	12	{ 14 1
Ordinary alcohol . . .	52.18	{ 34.78 13.04	46	24	{ 16 6
Methyl alcohol . . . .	87.50	{ 50.00 12.50	32	12	{ 16 4
Carbon bisulphide . . .	15.79	84.21	76	12	64
Chloroform . . . . .	10.04	{ 89.12 0.84	119.5	12	{ 106.5 1
Ether . . . . .	64.87	{ 21.62 13.51	74	48	{ 16 10

Here we find the smallest weight of carbon in any of these compounds is 12, and in the others it is some multiple of 12. This is true of all carbon compounds yet examined in this way; 12 is therefore adopted as the atomic weight of carbon.

The following data show why 14 is adopted as the atomic weight of nitrogen:—

Substance.	Per Cent of Nitrogen.	Per Cent of Other Elements.	Molecular Weight.	Weight of Nitrogen in Molecule.	Weight of Other Elements in Molecule.
Nitrous oxide . . . . .	63.64	36.36	44	28	16
Nitric oxide . . . . .	46.67	53.33	30	14	16
Cyanogen . . . . .	53.85	46.15	52	28	24
Cyanogen chloride . .	22.76	{ 57.73 19.51	61.5	14	{ 35.5 12
Hydrocyanic acid . . .	51.85	{ 44.44 3.71	27	14	{ 12 1
Ammonia . . . . .	82.35	17.65	17	14	3

This method is applicable only to elements which have gaseous compounds, or compounds that can, by heat, be converted into gases without decomposition.

There are many such compounds, yet some of the elements have no such compounds; hence, we can get no evidence in regard to their atomic weights by this method.

Again other elements have but one or two such compounds, so that the evidence furnished by them is by no means conclusive, for it might happen that the few volatile compounds of an element were not the simplest compounds of that element.

For example, suppose that olefiant gas, cyanogen, and ordinary alcohol were the only volatile compounds of car-

bon known, we should then be led to the false conclusion that 24 was the atomic weight of carbon.

Until within very recent times, however, this has been the only means available for determining the molecular weights of compounds, independently of the knowledge of the atomic weights.

Quite recently some new methods have been developed, which, while still more or less imperfect as far as methods of manipulation are concerned, give promise of being exceedingly useful to chemists.

These methods are based upon the behavior of dilute solutions; it would lead us too far to discuss them all here. One of them may, however, be briefly mentioned.

It has long been known that the freezing point of water is lowered by dissolving substances in it, as, for example, salt water does not so readily freeze as fresh water, so the solidifying point of any liquid is lowered by dissolving foreign substances in it.

**Raoult's Law.**—In 1887 François Raoult pointed out that, so long as the solution was dilute, there was a constant relation between the molecular weight of a substance and the freezing point of its solution, and that this fact might be made use of to determine the molecular weights of substances.

Raoult found that when weights of different substances which were proportional to their molecular weights were dissolved in the same amount of a solvent, they produced the same depression of the freezing point. This is known as Raoult's law, and may also be stated as follows: *When equal weights of different substances are dissolved in the same amounts of the same solvent, the lowerings of the freezing point are inversely proportional to the molecular weights of the dissolved substances.*

This law holds true only when the solution is dilute, and when there is no chemical action between the solvent and the substance dissolved.

With a delicate thermometer and suitable apparatus, it has been found possible to determine the lowering of the freezing point with considerable accuracy.

**Determination of Molecular Weights with the Aid of Raoult's Law.**—From Raoult's law, where the amounts of substance and solvent are constant, we have

$$M:m::l:L \text{ or } M=\frac{ml}{L},$$

where  $M$  and  $m$  are the molecular weights of different substances, and  $L$  and  $l$  are the corresponding lowerings of the freezing point produced by dissolving the same weights of the substances in the same amounts of the solvent.

The weights commonly used are 1 gram of substance in 100 grams of solvent.

Here, as in the case of Avogadro's theory, in order to find the molecular weight of a compound, it is necessary to know the molecular weight of one substance. One gram of the substance of known molecular weight is dissolved in 100 grams of the solvent and the lowering of the freezing point determined; the product of this lowering and the molecular weight—that is,  $ml$ —is a constant for all substances and needs to be determined but once. Let  $ml=K$ ; then the formula becomes—

$$M=\frac{K}{L}.$$

The constant  $K$  is different for different solvents, but always the same for the same solvent.

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$$M=\frac{K}{L}.$$

The constant  $K$  is different for different solvents, but always the same for the same solvent.

This simple and very satisfactory method for the determination of the molecular weights has been found of great

service with a very large class of compounds in organic chemistry, but does not seem to be applicable to the more common compounds of inorganic chemistry, for the probable reason that these undergo some fundamental changes when they dissolve.

### QUESTIONS IN REVIEW.

1. What is the theory of Avogadro? What led him to propose such a theory?

2. How does this theory lead to conclusions concerning the molecular weights of compounds?

3. What reasons are there for assuming that 36.5 is the molecular weight of hydrochloric acid?

4. Point out how the knowledge of the molecular weights of compounds aids in the determination of the atomic weights.

5. Knowing the specific gravity in the form of vapor, and the amount of sulphur contained in the following compounds to be as follows: —

	Per Cent of Sulphur.	Specific Gravity.
Sulphur dioxide . . . . .	50.00	2.222 (air = 1)
Sulphur trioxide . . . . .	40.00	40.000 (hydrogen = 1)
Hydrogen sulphide . . . .	94.12	1.180 (air = 1)
Carbon bisulphide . . . .	84.21	38.000 (hydrogen = 1)
Sulphuryl chloride . . . .	23.70	4.687 (air = 1)

what conclusions can you draw in regard to the atomic weight of sulphur?

6. What is Raoult's law? Show that the two statements of it given in the text express the same facts.

7. Suppose 1 gram of a substance the molecular weight of which was 90 were dissolved in 100 grams of acetic acid, and the lowering of the freezing point was found to be  $0.43^{\circ}$ , what would be the constant for acetic acid?

8. Using the constant determined in the previous problem, what would be the molecular weight of a substance which produced a lowering of the freezing point equal to  $0.26^{\circ}$  when one gram of it is dissolved in 100 grams of acetic acid?

## CHAPTER IX.

SPECIFIC HEAT. — LAW OF DULONG AND PETIT. — AID OF THIS LAW IN DETERMINATION OF ATOMIC WEIGHTS. — AID GAINED IN THE DETERMINATION OF ATOMIC WEIGHTS FROM THE PERIODIC LAW. — QUESTIONS IN REVIEW.

**Specific Heat.** — When equal weights of two different substances at the same temperature are subjected for the same length of time to the same amount of heat, it is found that they have different temperatures. If, for example, a kilogram of mercury and a kilogram of water are subjected to the same source of heat for ten minutes, at the end of the time the mercury will be very much hotter than the water, or, in other words, it requires more heat to raise the temperature of water through one degree than it does to raise the temperature of the same mass of mercury through one degree. The amount of heat required to raise the temperature of a definite weight of a substance, through a definite number of degrees, as compared with the amount of heat necessary to raise the temperature of the same amount of water through the same number of degrees, is called the *specific heat*. That is, when the specific heat of silver is given as 0.0570, it is meant that if the amount of heat required to raise the temperature of one gram of water  $1^{\circ}$  is 1, then the amount of heat required to raise the temperature of one gram of silver  $1^{\circ}$  is 0.0570. Or, to state it in a still different way, the quantity of heat that would raise the tem-

perature of any given weight of silver through  $1^{\circ}$  will raise the same weight of water through 0.0570.

**Law of Dulong and Petit.** — A knowledge of the specific heat of the elements has become of great importance in consequence of Dulong and Petit's discovery in 1819 of the law that *the product of the numbers representing the specific heat and the atomic weight of the elements is approximately a constant number*, as will be seen from the following table : —

	Specific Heat.	Atomic Weight.	Product.
Aluminium . . . . .	0.2143	27.04	5.79
Antimony . . . . .	0.0513	119.6	6.13
Arsenic . . . . .	0.0822	74.9	6.15
Bismuth . . . . .	0.0308	207.5	6.39
Cadmium . . . . .	0.0567	111.7	6.33
Calcium . . . . .	0.170	39.91	6.78
Cobalt . . . . .	0.1067	58.6	6.25
Copper . . . . .	0.0939	63.18	5.93
Gold . . . . .	0.0324	196.2	6.35
Iron . . . . .	0.1138	55.88	6.35
Lead . . . . .	0.0314	206.39	6.48
Magnesium . . . . .	0.2475	23.94	5.92
Mercury . . . . .	0.0332	199.8	6.63
Nickel . . . . .	0.1092	58.6	6.39
Platinum . . . . .	0.0324	194.37	6.29
Potassium . . . . .	0.1655	39.03	6.45
Silver . . . . .	0.0570	107.66	6.13
Sodium . . . . .	0.293	22.995	6.73
Sulphur . . . . .	0.178	31.98	5.69
Tin . . . . .	0.0555	117.35	6.51
Zinc . . . . .	0.0956	64.88	6.20

**Aid of this Law in Determination of Atomic Weights.** —

It will be seen that the product is not constant, varying,

as it does in the table given above, between 5.69 and 6.78. These variations may depend partly on errors incidental to the determination of the specific heats and atomic weights. It has been found that the specific heat of the same substance varies with the temperature at which it is determined. The specific heat of some elements, such as potassium, have been determined at temperatures very near their fusing points; others, like platinum, at temperatures far removed from their fusing points. A main cause, therefore, of the variations is doubtless the fact that all the determinations have not been made under similar physical conditions of the elements. Still other reasons exist which make it improbable that the numbers obtained should give a product that is exactly the same in each case.

Nevertheless, so long as the numbers obtained are approximately constant, the law is of direct help in the determination of the atomic weights.

We have  $\text{At. wt.} \times \text{Sp. ht.} = 6.25 \pm$ ;

hence 
$$\text{At. wt.} = \frac{6.25 \pm}{\text{Sp. ht.}}$$

Of course this will not give the exact atomic weight, but it comes to the aid of analysis; and by means of both, the atomic weight may be determined.

Let us take the case of silver, for example. Assuming that we know the atomic weight of chlorine to be 35.37, in silver chloride we find, by analysis, 107.66 parts of silver united with 35.37 parts of chlorine. If in silver chloride there is one atom of silver and one atom of chlorine, the atomic weight of silver would be 107.66; but if there were two atoms of silver to one of chlorine, the atomic weight of silver would be 53.83; if, on the other hand, there were two atoms of chlorine to one of silver, the atomic weight of

silver would be 215.32. Now it is impossible from analysis alone to determine which one of these, or other assumptions that might be made, is the correct one.

$$\text{From the relation, At. wt.} = \frac{6.25 \pm}{\text{Sp. ht.}},$$

$$\text{we obtain} \quad \text{At. wt. silver} = \frac{6.25}{0.057} = 109.6;$$

that is, we know that the atomic weight of silver is somewhere near 109.6; consulting the numbers obtained by analysis, 53.83, 107.66, 215.32, etc., we see at once that the atomic weight is 107.66.

The law of Dulong and Petit may also be stated as follows: *The specific heats of the elements are inversely proportional to their atomic weights.*

The discoverers themselves announced the law in still a different form; they said: *The atoms of all elements have exactly the same capacity for heat.* That is, to raise the temperature of an atom of any element  $1^{\circ}$ , the same quantity of heat is always necessary.

This law was proposed in 1819, but it was not generally adopted by chemists for a considerable time. One of the chief reasons for the delay was the fact that certain elements appeared to be exceptions to the law; that is, the product of their atomic weights and specific heats appeared to be very much less than 6.25, when the number obtained from a study of the gaseous compounds of the element was taken as the atomic weight.

For example, carbon, whose atomic weight as shown in its gaseous compounds is 12, has a specific heat of 0.132,

$$12 \times 0.132 = 1.58;$$

the product 1.58 is very far from 6.25.



In a similar manner the elements silicon and boron gave numbers much smaller than 6.

Later, however, it was observed that if the specific heats were taken at high temperatures, even these elements were not exceptions to the law.

To be strictly true, then, the law needs some modification. It might be more correctly stated as follows :—

The specific heats of the elements vary with the temperature ; but for every element there is a temperature above which the variation is very slight, and the product of the atomic weight and of the value of the specific heat, as determined above this temperature, is approximately constant.

The methods for the determination of the atomic weights given in the last chapter and in this, have been the chief thoroughly reliable ones available, and the ones which have been made use of in determining the atomic weights we now use. When both methods are applied to the same element, they lead to the same atomic weight.

There is, however, another very strong proof that the numbers we now have and call the atomic weights have some characteristic and fundamental relation to the elements to which they refer.

Almost as soon as the present system of atomic weights was extended to the elements, a very close relation was discovered between the properties of an element and its atomic weight ; that is, if the atomic weight of an element is known, many of its properties can be predicted without farther study. This relation is shown in what is termed the Periodic Law of the elements. If the elements, omitting hydrogen, are written in the order of their atomic weights, beginning with the lowest, and ending with the highest, it will be found that at regular periods elements with similar properties recur.

The eighth element is like the first ; the ninth element is

like the second ; the tenth is like the third ; the fifteenth is like the first and eighth ; the sixteenth like the second and ninth ; the seventeenth like the third and tenth ; and so on.

**Aid gained in the Determination of Atomic Weights from the Periodic Law.** — It is not desired here to discuss the periodic law, which is, perhaps, the most far-reaching discovery in chemistry of modern times, but which will be better understood after the student has a more extended acquaintance with the properties of the elements and their compounds. It is introduced here merely for the sake of showing that our atomic weights determined by the methods described really are characteristic of the elements, and therefore rest on a firm basis of facts ; also to point out that the periodic law itself is an aid in the determination of the atomic weights of new elements. The properties of an element at once show us to what other elements it is similar ; and, therefore, where it belongs in the periodic system of elements ; and thus, too, between what numbers its atomic weight must fall.

For example, take the recently discovered element germanium. By analysis its oxide was found to contain 36 parts of germanium to 16 parts of oxygen. Now if the oxide contained two atoms of germanium and one of oxygen, the atomic weight of germanium would be 18 ; if there were one atom of each element present, then the atomic weight of germanium would be 36, and so on. By analysis, we know that the atomic weight of germanium is 18, or some multiple of 18. A study of the chemical properties of germanium showed that it belongs in the periodic system between the elements gallium, with an atomic weight of 70, and arsenic, with an atomic weight of 75 ; that is, according to the periodic law, germanium must have an atomic weight

between 70 and 75, and according to analysis it must be some multiple of 18; it must therefore be 72 ( $18 \times 4$ ).

### QUESTIONS IN REVIEW.

1. What is meant by specific heat?
2. What is the law of Dulong and Petit? Convince yourself that the three statements of the law given in this chapter are really different ways of expressing the same fact.
3. How does the law of Dulong and Petit aid in the determination of the atomic weights of the elements?
4. Pure cadmium chloride contains 61.225 % cadmium and 38.775% chlorine. Assuming that the atomic weight of chlorine is 35.37, and the specific heat of cadmium is 0.0567, what is the atomic weight of cadmium?
5. How does the periodic law furnish any evidence in regard to the atomic weights?

## CHAPTER X.

DETERMINATION OF FORMULÆ. — VALENCE. — REPLACING POWER. — EQUIVALENT WEIGHT. — RADICALS. — OXIDATION AND REDUCTION. — QUESTIONS IN REVIEW.

**Determination of Formulæ.** — When we know the atomic weights of the elements, it becomes a simple matter to determine the formulæ of chemical compounds.

In the first place, since the formula indicates the molecular weight and the composition of the compound, it is necessary that these should be determined by experiment. The composition is determined by making a quantitative analysis of the compound; the molecular weight may be determined by one of the methods described in Chapter VIII.; still other methods are occasionally used. Knowing the molecular weight and the percentage composition of the compound, the weight of each element present in the molecule is readily calculated; also, knowing the atomic weight of each element, it is a simple matter to determine how many atoms are required to make up the weight of that element present in the molecule, and to write the formula accordingly. For example, the percentage composition of ordinary alcohol is

Carbon . . . . .	52.18 %
Hydrogen . . . . .	13.04 %
Oxygen . . . . .	34.78 %
Total . . . . .	<u>100.00 %</u>

The specific gravity of alcohol vapor is 23 (hydrogen=1) ; its molecular weight is therefore  $(23 \times 2 = 46)$  46.

52.17 % of 46 is 24.00, the weight of carbon in the molecule.

13.04 % of 46 is 5.998, or 6.00, the weight of hydrogen in the molecule.

34.78 % of 46 is 15.9988, or 16.00, the weight of oxygen in the molecule.

That is, the molecule of alcohol, which weighs 46 times as much as an atom of hydrogen, contains

24 parts carbon,
6 parts hydrogen,
16 parts oxygen,
46 parts alcohol.

But since one atom of carbon weighs 12,

*two* atoms will weigh 24 ;

one atom of hydrogen weighs 1,

*six* atoms will weigh 6 ;

one atom of oxygen weighs 16,

*one* atom of oxygen is present in the molecule of alcohol and its formula will be  $C_2H_6O$ .

Again, supposing the molecular weight of sulphuric acid to be 98 and its composition to be

Hydrogen . . . . .	2.04 %
Sulphur . . . . .	32.65 %
Oxygen . . . . .	65.31 %
Total . . . . .	100.00 %

The atomic weight of hydrogen being 1, sulphur 32, and oxygen 16, we have

2.04 % of 98 is 1.999, or  $2.00, \frac{2.00}{1} = 2$  atoms of H.

32.65 % of 98 is 31.99, or 32.00,  $\frac{32}{32} = 1$  atom of S.

65.31 % of 98 is 64.00, or  $\frac{64}{16} = 4$  atoms of O.

The formula for sulphuric acid is therefore  $\text{H}_2\text{SO}_4$ .

In some cases it has not been possible to determine the molecular weight of the compound. In such cases it is customary to give the compound the *simplest* formula that will truly express its composition.

Sodium chloride (common salt) is such a compound ; we know its composition to be

Sodium . . . . .	39.31 %
Chloride . . . . .	60.69 %
Total . . . . .	<u>100.00 %</u>

but we do not know its molecular weight.

Let us assume for a moment that the molecular weight of sodium chloride is 100 ; then the amount of sodium in the molecule will be 39.31 and the chlorine in the molecule will be 60.69.

Since the atomic weight of sodium is 23 and chlorine 35.5, the number of atoms of each element present in the molecule, if 100 is the molecular weight of sodium chloride, is

$$\frac{39.31}{23} = 1.70 \text{ atoms of sodium ;}$$

$$\frac{60.69}{35.5} = 1.70 \text{ atoms of chlorine.}$$

That is, if 100 is the molecular weight of sodium chloride, there would be 1.7 atoms of sodium and 1.7 atoms of chlorine in the molecule of sodium chloride.

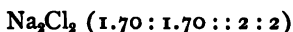
But according to our definition of atoms, they are indivisible, and to speak of fractions of an atom is, therefore, absurd.

We do know, however, that in the molecule of salt, for every 1.7 atoms of sodium present, there must be 1.7 atoms of chlorine; or, what is the same thing, for every atom of sodium, there must be one atom of chlorine; hence the simplest formula is NaCl.

Or, in general, the numbers obtained by dividing the percentage of the element present in the compound by the atomic weight of that element do not represent the actual number of atoms present in the molecule; but the numbers thus obtained have the same ratio to one another that the actual numbers have. Where the molecular weight of the compound is unknown, it is customary to take the simplest whole numbers which have this ratio. In the case of sodium chloride just considered, we have

$$1.70 : 1.70 :: 1 : 1 ;$$

the formula, then, is NaCl. Of course the formula



would just as truly express the composition of sodium chloride, or  $\text{Na}_x\text{Cl}_x$ , where  $x$  is any whole number, would also express the composition; but so long as there is no way of deciding which is correct, the simplest formula is used.

We do not know the molecular weight of calcium chloride, but analysis of the dried salt gives

Calcium . . . . .	36.04 %
Chlorine . . . . .	63.96 %
Total . . . . .	100.00 %

Since 40 is the atomic weight of calcium and 35.5 is the atomic weight of chlorine, we have

$$\frac{36.04}{40} = 0.900, \quad \frac{63.96}{35.5} = 1.800,$$

and  $0.900 : 1.800 :: 1 : 2$  ;

the formula for calcium chloride is  $\text{CaCl}_2$ .

Acetic acid has the composition

Carbon . . . . .	40.00 %
Hydrogen . . . . .	6.60 %
Oxygen . . . . .	53.40 %
Total . . . . .	100.00 %

$$\frac{40}{12} = 3.3 ; \quad \frac{6.6}{1} = 6.6 ; \quad \frac{53.4}{16} = 3.3 ;$$

$$3.3 : 6.6 : 3.3 :: 1 : 2 : 1.$$

The simplest formula for acetic acid is, therefore,  $\text{CH}_2\text{O}$ .

Now if  $\text{CH}_2\text{O}$  were the true formula for acetic acid, its molecular weight would be 30 ( $12 + 2 + 16$ ). But the molecular weight of acetic acid has been found to be 60 ; its true formula is, therefore, twice  $\text{CH}_2\text{O}$  or  $\text{C}_2\text{H}_4\text{O}_2$ . Again, lactic acid has been found to have exactly the same composition that acetic acid has ; hence its simplest formula would be  $\text{CH}_2\text{O}$ . But its molecular weight has been found to be 90 ; hence its true formula is  $3(\text{CH}_2\text{O})$ , or  $\text{C}_3\text{H}_6\text{O}_3$ .

**Valence.**— In Exp. 66 we have seen that one volume of chlorine unites with one volume of hydrogen. In Exp. 26 we learn that one volume of oxygen unites with two volumes of hydrogen. Exps. 81 and 82 show that one volume of nitrogen combines with three volumes of hydrogen. We



thus see that elements differ in their power to combine with hydrogen. If we compare the formulæ of compounds which have been calculated by the methods just outlined, we see the same characteristic property. Taking first the formulæ of some of the more common hydrogen compounds, we have

I.	II.	III.	IV.
FlH	OH <sub>2</sub>	NH <sub>3</sub>	CH <sub>4</sub>
ClH	SH <sub>2</sub>	PH <sub>3</sub>	SiH <sub>4</sub>
BrH	SeH <sub>2</sub>	AsH <sub>3</sub>	
IH	TeH <sub>2</sub>	SbH <sub>3</sub>	

From these formulæ we see that the elements which combine with hydrogen may be divided into four classes, according as one atom of the element combines

- (1) with one atom of hydrogen,
- (2) with two atoms of hydrogen,
- (3) with three atoms of hydrogen,
- (4) with four atoms of hydrogen.

It has been found that every element which combines with hydrogen belongs to one of these four classes.

If we were to classify the elements in a similar manner, according to their ability to combine with chlorine, we should find that here, too, they fall into classes as in the case of hydrogen; only, in the case of chlorine, the number of classes is still greater, as is shown by the following formulæ:—

I.	II.	III.	IV.	V.	VI.
LiCl	MgCl <sub>2</sub>	PCl <sub>3</sub>	CCl <sub>4</sub>	PCl <sub>5</sub>	WCl <sub>6</sub>
NaCl	CaCl <sub>2</sub>	AsCl <sub>3</sub>	SiCl <sub>4</sub>	AsCl <sub>5</sub>	—
KCl	ZnCl <sub>2</sub>	SbCl <sub>3</sub>	SnCl <sub>4</sub>	SbCl <sub>5</sub>	—

The elements exhibit a still greater diversity in their ability to combine with oxygen, as is made clear by the following formulæ : —

I.	II.	III.	IV.
$\text{Na}_2\text{O}$	$\text{MgO}$	$\text{Fe}_2\text{O}_3$	$\text{CO}_2$
	$(\text{Mg}_2\text{O}_2)^1$		$(\text{C}_2\text{O}_4)$
$\text{K}_2\text{O}$	$\text{CaO}$	$\text{P}_2\text{O}_5$	$\text{SiO}_2$
	$(\text{Ca}_2\text{O}_2)$		$(\text{Si}_2\text{O}_4)$
$\text{N}_2\text{O}$	$\text{NO}$	$\text{N}_2\text{O}_3$	$\text{NO}_2$
	$(\text{N}_2\text{O}_2)$		$(\text{N}_2\text{O}_4)$
V.	VI.	VII.	VIII.
$\text{P}_2\text{O}_5$	$\text{SO}_3$	$\text{Cl}_2\text{O}_7$	$\text{RuO}_4$
	$(\text{S}_2\text{O}_6)$		$(\text{Ru}_2\text{O}_8)$
$\text{As}_2\text{O}_5$	$\text{SeO}_3$	$\text{Mn}_2\text{O}_7$	$\text{OsO}_4$
	$(\text{Se}_2\text{O}_6)$		$(\text{Os}_2\text{O}_8)$
$\text{N}_2\text{O}_5$	$\text{TeO}_3$		
	$(\text{Te}_2\text{O}_6)$		

An inspection of these tables shows that every atom of an element has an inherent power of holding in combination a certain number of other atoms, this number depending upon the nature of the atoms held in combination.

This power or property of the atoms is called *valence*. Those atoms which unite with one atom of hydrogen, such as chlorine, bromine, and iodine, are said to be *univalent*; those atoms which unite with two atoms of hydrogen, such as oxygen and sulphur, are said to be *bivalent*; similarly, the

<sup>1</sup> Since in groups I., III., V., and VII. we have the number of atoms of oxygen which combine with *two* atoms of the elements, the formulæ given in groups II., IV., VI., and VIII. are doubled in the parentheses so as to give the number of atoms of oxygen combined with *two* atoms of the element in these cases also, for comparison with the other groups.

atoms of nitrogen, arsenic, etc., are *trivalent*, and the atoms of carbon and silicon are *quadrivalent*.<sup>1</sup>

Many elements do not unite with hydrogen: the valence of such elements is determined by the number of atoms of some other univalent element which one atom of the element in question will unite with. For example, silver does not unite directly with hydrogen, but it is considered univalent because one atom of silver unites with one atom of univalent chlorine, as shown in the formula,  $\text{AgCl}$ , for silver chloride.

For the same reason the elements calcium, magnesium, and zinc are *bivalent*; phosphorus, arsenic, and antimony are both *trivalent* and *pentavalent*.

Since one atom of oxygen combines with two atoms of hydrogen in water,  $\text{H}_2\text{O}$ , it is considered bivalent; it will therefore combine with two atoms of any univalent element, such as  $\text{K}_2\text{O}$  in potassium oxide, or  $\text{Na}_2\text{O}$  in sodium oxide, or  $\text{KOH}$  in potassium hydroxide. Magnesium, calcium, and zinc are bivalent, since they each combine with two atoms of chlorine. In their oxides,  $\text{MgO}$ ,  $\text{CaO}$ , and  $\text{ZnO}$ , they are still considered bivalent, since one atom of the element combines with one atom of the bivalent oxygen.

For the same reason, when one atom of an element combines with two atoms of oxygen, it is regarded as quadrivalent; and when one atom of any element holds in combination three atoms of oxygen, it is regarded as hexivalent towards oxygen.

Formerly it was believed that each element had but one unchangeable valence.

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<sup>1</sup> Instead of *valence* the term *quantivalence* is sometimes used. *Atomicity* is frequently met with as synonymous with *valence*, and the adjectives *monatomic*, *diatomic*, *triatomic*, and *tetratomic* are used instead of univalent, bivalent, trivalent, and quadrivalent.

The theory of constant valence very well illustrates the evil effect a false theory may have upon a science.

At first, the few compounds known where the same element appeared to have more than one valence — such as,

	phosphorus trichloride, $\text{PCl}_3$ ;
and	phosphorus pentachloride, $\text{PCl}_5$ ;
	sulphur bichloride, $\text{SCl}_2$ ;
and	sulphur tetrachloride, $\text{SCl}_4$ —

were explained by various assumptions, which were not warranted by any of the properties of the compounds themselves.

Many compounds have undoubtedly been neglected, or set aside, without thorough study, because the formulæ which expressed their composition could not be written in accordance with the preconceived notions of the valence of the elements composing them. Finally, the evidence against the conception of constant valence accumulated to such an extent that it could no longer be ignored.

The evidence, at first, seemed to indicate that when the valence varied, it was either increased or decreased by 2, so that if the valence were expressed by an odd number in one case, it must always be expressed by an odd number; or, if it were even in one instance, it must always be even; that is, if the valence of an element were 1 it might also be 3 or even 5, but could not be either 2 or 4, or if it were 2 it might also be 4 or 6, but could not be 1, 3, or 5.

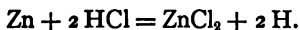
Hence there gradually arose the classification of the elements into two classes; (1) *artiads* or those elements which had an even valence; and (2) *perissads* or those elements which had an odd valence. This classification, which was used in most text-books on chemistry a few years ago, is still met with occasionally.

The idea that the valence is constant, or that it varies by 2, is no longer generally held, but has given way to less definite and more elastic views concerning this subject.

The valence of nitrogen, for example, appears to vary from 1 to 5, as exhibited by the following formulæ:  $N_2O$ ,  $NO$ ,  $NH_3$ ,  $N_2O_3$ ,  $NO_2$ ,  $N_2O_5$ .

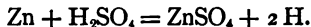
Nevertheless, the idea of valence will be of great help to the student in his effort to become familiar with chemical formulæ; for in the compounds more commonly met with the elements usually exhibit a constant valence, or not more than two different valences, so that a knowledge of these will enable him to write the formulæ of the simple compounds without difficulty.

**Replacing Power; Equivalent Weight.** — Zinc being a bivalent metal combines with two atoms of chlorine. When hydrochloric acid acts upon zinc, in order that the zinc may get its two atoms of chlorine, it must have two molecules of hydrochloric acid to act upon; thus, —



From this we see that one atom of zinc drives out, or replaces, two atoms of hydrogen in hydrochloric acid.

The molecule of sulphuric acid contains two atoms of hydrogen; one atom of zinc would replace both these, and zinc sulphate would have the formula  $ZnSO_4$ .



In phosphoric acid,  $H_3PO_4$ , there are three atoms of hydrogen; the formula of zinc phosphate, in which all the hydrogen has been replaced by zinc, is  $Zn_3(PO_4)_2$ , for three zinc atoms would replace six hydrogen atoms; to furnish six hydrogen atoms, two molecules of phosphoric acid are

required. The student should bear in mind that if the normal zinc phosphate were analyzed, and its formula calculated from the analysis as already described, the same formula for zinc phosphate would be reached.

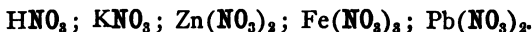
Since one atom of zinc is capable of replacing two atoms of hydrogen, and since 65 is the atomic weight of zinc, it follows that 65 parts, by weight, of zinc, replace, or are equivalent to, 2 parts, by weight, of hydrogen; or, that 32.5 parts of zinc are equivalent to 1 part hydrogen. (Compare this with the result obtained in Exp. 51.)

In general, the replacing power of an element for hydrogen is the same as its valence; it follows, therefore, that the equivalent weight of an element is equal to its atomic weight divided by its valence.

**Radicals.**—An inspection of the formulæ of compounds shows that in many cases certain groups of elements are present in many compounds. For example, in sulphuric acid and in all the sulphates we have the group  $\text{SO}_4$ ; thus,—



In nitric acid and nitrates there is the group  $\text{NO}_3$ , as—



In ammonium salts is the group  $\text{NH}_4$ , as—



Such groups, which pass from compound to compound apparently unchanged, and in many chemical changes act much as an element, are frequently spoken of as *radicals*. The radicals, like the elements, may be considered as *univalent*, *bivalent*, *trivalent*, etc., according as they unite with one, two, three, etc., hydrogen atoms or other univalent atoms; thus  $\text{NO}_3$  is univalent, because it unites with one

atom of hydrogen,  $\text{HNO}_3$ ;  $\text{NH}_4$  is univalent, because it unites with one atom of chlorine,  $\text{NH}_4\text{Cl}$ , or with a univalent radical, as  $\text{NO}_3$  in  $\text{NH}_4\text{NO}_3$ ;  $\text{SO}_4$  is bivalent, because it unites with two atoms of hydrogen in  $\text{H}_2\text{SO}_4$ ;  $\text{PO}_4$  is trivalent, because it unites with three atoms of hydrogen in  $\text{H}_3\text{PO}_4$ .

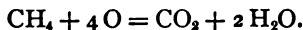
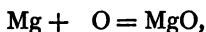
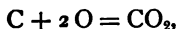
The valence of the radical is, however, believed to be a direct consequence of the valence of the elements which compose it; in  $\text{NH}_4$ , for example, the valence of nitrogen is believed to be 5, and since it is united with only four atoms of hydrogen, it is capable of uniting with one more univalent atom, so that the group or radical  $\text{NH}_4$  appears to have a valence of 1.

These radicals have no actual existence; they cannot be separated from their compounds; no substances with the formulæ  $\text{NO}_3$ ,  $\text{NH}_4$ ,  $\text{SO}_4$ ,  $\text{PO}_4$ , etc., have ever been prepared. They exist only in combination with other elements.

The formulæ of compounds, besides expressing the facts already referred to in Chapter V., also express the valence of the elements composing the compound with more or less clearness.

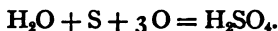
The student should, in using the formulæ, observe the valence of the different elements and radicals composing them, for these will be of assistance in remembering and writing the formulæ of different chemical compounds.

**Oxidation and Reduction.**—When any element or any substance combines with oxygen, it is said to be oxidized; the process is referred to as a process of oxidation. The following equations represent processes of oxidation:—

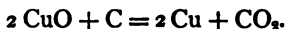


Compounds which easily give up their oxygen to other substances are called *oxidizing* agents. Thus, copper oxide,  $\text{CuO}$ , in Exp. 58, gives up its oxygen to the hydrogen; the copper oxide is, therefore, spoken of as an oxidizing agent.

Nitric acid,  $\text{HNO}_3$ , is said to be an oxidizing agent because it breaks up, giving a part of its oxygen to other things. Thus in Exp. 161 the sulphur was converted into sulphuric acid by treatment with nitric acid.



On the other hand, the extraction of oxygen from a substance is called *reduction*. Any substance which has the power of combining with oxygen and taking it away from other things is called a *reducing* agent. Carbon is a reducing agent in Exp. 107.

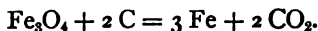
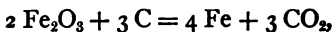


Hydrogen, in Exp. 58, is also a reducing agent.



We speak of the *reduction* of the metals from their ores. The ores are frequently the oxides of the metals, and they are smelted with carbon (coal, coke, or charcoal) and the oxygen taken away, the metal remaining in a state of greater or less purity.

For example, the reduction of iron from its ores may be represented by the following equations, since the chief ores of iron are the oxides  $\text{Fe}_2\text{O}_3$  and  $\text{Fe}_3\text{O}_4$ :—

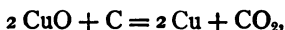


The actual reactions have been found to differ in the different parts of the blast furnace, and are not quite as simple



as here expressed, yet the reactions are of this general character.

The processes of oxidation and reduction are seen to be complementary. In the equation



the copper is reduced, and the carbon is oxidized. Whenever we speak of a substance as having been oxidized, it implies another substance which has furnished the oxygen, and in consequence has itself been reduced. When we speak of a substance as having been reduced, it implies another substance which has taken away the oxygen and has in consequence been oxidized.

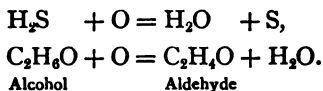
**Oxidizing and Reducing Flames.**—We can now readily understand the distinction between the *oxidizing* and *reducing* flames.

If a substance is so placed in a flame that it may, while it is heated, come in contact with the outside air, it will tend to combine with the oxygen of the air; it will be *oxidized*. That part of the flame which fulfils these conditions—that is, the outer part and especially the upper part of the flame—is called the *oxidizing* flame.

On the other hand, if the substance is held well within the flame, so that it is completely surrounded by the flame, the air cannot come in contact with it. Since the substance is completely surrounded with hot carbon and unburned gases, which will be burned if they can get oxygen, the substance, if it contains oxygen, will tend to give it up to these hot gases, or, in other words, the substance will tend to be reduced. The part of the flame that fulfils these conditions—that is, the inner part—is called the *reducing* flame.

It frequently happens that when substances which contain hydrogen are oxidized, the oxygen, instead of uniting with

the substance, simply removes a part of the hydrogen, forming with it water. Thus



The change here has been to remove a part or all of the hydrogen from the compound ; it is spoken of as oxidation. We may speak of oxidizing hydrogen sulphide to sulphur or of oxidizing alcohol to aldehyde, although no oxygen has been added to the compound in either case.

Similarly, when hydrogen is added to a compound, even without the abstraction of oxygen, the process is called reduction ; that is, we speak of *reducing* sulphur to hydrogen sulphide, or of *reducing* aldehyde to alcohol, although no oxygen has been removed from the compound in either case.

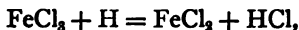
An examination of the tables given in the early part of this chapter will show that where elements have more than one valence, they exhibit their minimum valence towards hydrogen and their maximum valence towards oxygen.

Towards hydrogen, for example, sulphur never exhibits a valence greater than two, as in  $\text{SH}_2$ , while towards oxygen it has a valence of six, as shown in  $\text{SO}_3$ . Phosphorus exhibits a valence of three towards hydrogen,  $\text{PH}_3$ , and a valence of five towards oxygen,  $\text{P}_2\text{O}_5$ . So it happens that when an element is oxidized its valence is increased, and when it is reduced its valence is diminished. From this fact the chemist has come to look upon any change which increases the valence of an element as oxidation, whether or not oxygen has been added ; and, conversely, any change in which the valence of the element is diminished is looked upon as a reduction, whether oxygen has been abstracted or not. Thus

the chemist speaks of oxidizing ferrous chloride to ferric chloride,



although there has been no addition of oxygen. The conversion of ferric chloride to ferrous chloride,



is spoken of as reduction, although there has been no oxygen removed from the compound. Likewise the conversion of mercurous compounds into mercuric compounds, cuprous compounds into cupric compounds, stannous compounds into stannic compounds, by whatever process it may be performed, is called *oxidation*, and the reverse process is called *reduction*.

By *oxidation*, then, the chemist understands a process which adds oxygen to a substance, or one which abstracts hydrogen, or one which converts the compound in which the element exhibits a low valence into a compound in which the element exhibits a higher valence.

By *reduction* the chemist understands a process which abstracts oxygen from a substance, or one which adds hydrogen to it, or which changes a compound in which a given element has a high valence to another compound in which the same element has a lower valence.

#### QUESTIONS IN REVIEW.

1. One of the sulphides of phosphorus has the composition —

Phosphorus, P . . . . .	27.93%
Sulphur, S . . . . .	72.07%
Total . . . . .	100.00%

The specific gravity of its vapor is 111 (hydrogen = 1). Atomic weight of P is 31; of S, 32. What is its formula?

## 2. The composition of chloroform is —

Carbon, C . . . . .	10.04 %
Hydrogen, H . . . . .	0.84 %
Chlorine, Cl . . . . .	89.12 %
Total . . . . .	100.00 %

The specific gravity of its vapor is 4.150 (air = 1). Atomic weights are C = 12, H = 1, Cl = 35.5. What is the formula for chloroform?

## 3. Ferric chloride has the composition —

Iron, Fe . . . . .	34.46 %
Chlorine, Cl . . . . .	65.54 %
Total . . . . .	100.00 %

Its molecular weight is in doubt, the atomic weight of iron being 56, and chlorine, 35.5. What is the simplest formula for ferric chloride?

## 4. Crystallized barium chloride contains —

Barium, Ba . . . . .	56.15 %
Chlorine, Cl . . . . .	29.09 %
Water, H <sub>2</sub> O . . . . .	14.76 %
Total . . . . .	100.00 %

Its molecular weight is unknown. The atomic weight of barium is 137; chlorine, 35.5; and the molecular weight of water is 18. What is the simplest formula for crystallized barium chloride?

5. What is meant by valence? How is the valence of elements determined?

6. How does a knowledge of the valence of the elements help in reaching conclusions concerning the formulæ of compounds?

7. What is meant by the equivalent weight of an element? (See Exp. 51.) What relation has the equivalent weight of an element to its atomic weight?

8. What is meant by the term *radical* as employed in chemistry.

9. Review the formulæ given in Chapter VI., and from these satisfy yourself as to the valence of the different radicals and elements in these compounds.

10. Define fully the terms *oxidation* and *reduction* as employed by chemists.

11. Explain fully the differences between the oxidizing and reducing flames.

## CHAPTER XI.

GAY-LUSSAC'S GAS VOLUME LAW. — AVOGADRO'S THEORY. —  
ATOMS AND MOLECULES OF ELEMENTS. — NASCENT CONDI-  
TION. — QUESTIONS IN REVIEW.

**Gay-Lussac's Gas Volume Law.** — We have learned that one volume of chlorine unites with one volume of hydrogen to form two volumes of hydrochloric acid ; that one volume of oxygen unites with two volumes of hydrogen to form two volumes of water vapor ; that two volumes of ammonia gas contain one volume of nitrogen and three volumes of hydrogen ; that two volumes of nitrous oxide contain one volume of oxygen and two volumes of nitrogen ; and that two volumes of nitric oxide contain one volume of oxygen and one volume of nitrogen.

In 1805 Gay-Lussac and Humboldt determined the volumes of hydrogen and oxygen that could be obtained from a definite volume of water vapor. The simple volume relations observed here led Gay-Lussac to investigate other similar cases, and in 1808 he announced his gas volume law, which may be stated as follows : —

*When two or more gaseous substances combine chemically, the volumes of each bear a simple relation to one another ; if the product formed is gaseous, its volume also bears a simple relation to the volumes of the constituents.*

If this law holds true, and if we adopt the atomic theory, which assumes that chemical action is caused by the union

of the atoms with one another in the simplest numerical proportions, it follows that —

*The number of atoms contained in a given volume of a gaseous substance bears a simple relation to the number of atoms contained in the same volume of any other gas.*

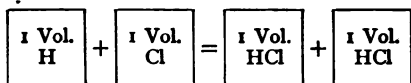
**Avogadro's Theory.** — It will be seen that this statement is closely related to the theory that Avogadro put forward, three years later, in 1811, that —

*Equal volumes of all gases, under like conditions of temperature and pressure, contain the same number of molecules.*

**Atoms and Molecules of Elements.** — Avogadro's theory was not, however, generally accepted at that time. One of the reasons for this was that chemists did not distinguish between the *atom* and the *molecule* of the elementary gases.

If this theory is true, such a distinction must be made, as a moment's consideration will make clear.

We know by experiment that —

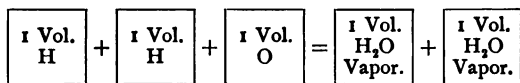


Let us suppose that the one volume of hydrogen contains  $n$  molecules of hydrogen, then, according to Avogadro's theory, the volume of chlorine also contains  $n$  molecules of chlorine, and each volume of hydrochloric acid contains  $n$  molecules of hydrochloric acid; there are therefore  $2n$  molecules of hydrochloric acid. That is,  $n$  molecules of hydrogen and  $n$  molecules of chlorine give  $2n$  molecules of hydrochloric acid.

But we know that every molecule of hydrochloric acid contains at least one atom of hydrogen and one atom

of chlorine. There must therefore be at least  $2n$  atoms of hydrogen in  $n$  molecules of hydrogen, and  $2n$  atoms of chlorine in  $n$  molecules of chlorine. Hence each molecule of hydrogen contains at least two atoms of hydrogen, and each molecule of chlorine contains at least two atoms of chlorine.

A similar study of the volume relations of hydrogen and oxygen and the water vapor formed by their union indicates that there are at least two atoms in the molecule of hydrogen, and also that there are at least two atoms in the molecule of oxygen, as is seen from the following consideration:—



That is,  $2n$  molecules of hydrogen and  $n$  molecules of oxygen give  $2n$  molecules of water vapor.

Since each molecule of water contains at least two atoms of hydrogen and one atom of oxygen, it follows that  $4n$  atoms of hydrogen and  $2n$  atoms of oxygen have come from  $2n$  molecules of hydrogen and  $n$  molecules of oxygen. Or, in other words, that one molecule of hydrogen contains at least two atoms of hydrogen, and that one molecule of oxygen contains at least two atoms of oxygen.<sup>1</sup> Every gaseous compound of hydrogen when considered in this way points to the presence of two atoms of hydrogen in the molecule of hydrogen. Nitrogen, chlorine, oxygen,—in fact, all the

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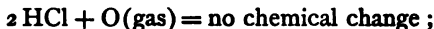
<sup>1</sup> The student should observe that it is possible for these volume relations to exist if there were any even number of atoms in the molecule greater than two, but the assumption of two atoms in the molecule gives a wholly satisfactory and much the simplest explanation of the phenomena observed.

elementary gases, for the same reasons,—are thought to have two atoms in their molecules.

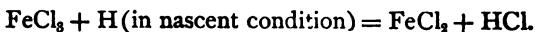
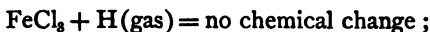
The question at once arises: Is there anything in the properties of the gases themselves which would indicate that the atom and the molecule are not the same? This question must be answered in the affirmative. The phenomena of the *nascent* state are best explained by the assumption that the molecule and the atom of the element are not the same.

**Nascent Condition** (from Latin *nascens* “being born”).—An element is said to be in the *nascent state* or *nascent condition* at just the instant it is set free from its compounds. Many of the elements are very much more active, chemically, the instant they are set free from their compounds than at any other time.

Oxygen gas, for example, may be passed through a solution of hydrochloric acid without causing any chemical change in the hydrochloric acid; but if we put into the hydrochloric acid solution a substance which is setting oxygen free, the oxygen the moment it is set free can take the hydrogen away from the chlorine; thus,—



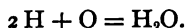
Again, hydrogen gas passed into a solution of ferric chloride produces no chemical change. But if we put into the ferric chloride a little zinc and hydrochloric acid, so that the hydrogen in the nascent condition comes in contact with the ferric chloride, it is converted into ferrous chloride; thus,—



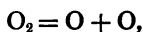


Many other examples of this peculiar activity of elements in the nascent condition are known. The simplest, and in fact the only explanation that chemists have of this phenomenon, is that the instant the element is set free, it is in the *atomic* condition, and that the atoms are very ready to combine with other atoms for which they have an attraction. But if such other atoms are not present, they unite with atoms of their own kind to form molecules of the element, and thereby lose much of their activity. For before they can act after they are combined to form molecules, some force must be employed to break up these molecules.

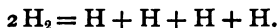
Thus nascent hydrogen and nascent oxygen combine directly with one another to form water.



But we know that the two gases do not combine without the aid of heat; that is, before they can combine, the molecule of hydrogen and the molecule of oxygen must be broken up into their atoms;—



and



We have seen (in Exps. 91 and 92) that nitrous oxide,  $\text{N}_2\text{O}$ , supports combustion quite as well as pure oxygen; yet when a substance burns in nitrous oxide, it unites only with the oxygen and leaves the nitrogen in the free condition. A mixture of nitrogen and oxygen containing the same proportions of the two gases as are present in nitrous oxide does not support combustion nearly so well as the nitrous oxide does. This is usually explained by the assumption that the molecule of nitrous oxide,  $\text{N}_2\text{O}$ , is more

easily broken up than the molecule of oxygen,  $O_2$ , and hence the oxygen is more easily obtained in the atomic condition from the nitrous oxide than it is from the oxygen gas.

Again, the chemist has no explanation of the phenomena of *allotropism* (see Exp. 104), except to assume that the molecule of the element in its different forms contains a different number of atoms.

Certain mathematical deductions made from a study of the specific heat of the vapors of some of the elements under a constant pressure, and the specific heat of the same vapor determined under a constant volume, have led to a similar conclusion; *i.e.* that some of the elements have more than one atom in the molecule.

The evidence is therefore sufficient to assume that the molecules of the elementary gases each contain two atoms.

This enables us to see clearly why the molecular weight of any gas is twice its specific gravity referred to hydrogen.

For since the molecular weight of hydrogen itself is 2, — that is, twice its specific gravity (hydrogen = 1), — it follows from Avogadro's theory (see page 142) that the molecular weight of any gas is twice its specific gravity referred to hydrogen.

The following proportions may make this statement a little clearer: —

$$\frac{\text{Molecular weight of any gas}}{\text{Molecular weight of hydrogen}} = \frac{\text{Sp. gr. of any gas}}{\text{Sp. gr. of hydrogen}},$$

whence

$$\text{M. wt. of any gas} = \frac{\text{Sp. gr. any gas} \times \text{M. wt. of H} (= 2)}{\text{Sp. gr. H} (= 1)}.$$

Since the elementary gases have two atoms in the molecule, it follows that their molecular weights are twice their

atomic weight, and that their specific gravities referred to hydrogen are the same as their atomic weights.

Not all the elements have two atoms in the molecule, however. Many of the metals are believed to have but one. The specific gravity of mercury vapor is such that its molecular weight and atomic weight appear to be identical.

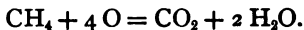
The specific gravity of sulphur vapor, just above its boiling point, indicates that its molecular weight is 192, but its atomic weight is believed to be 32; hence there must be six ( $\frac{192}{32} = 6$ ) atoms of sulphur in the molecule of sulphur.

As the temperature is increased, the specific gravity of sulphur vapor grows less and less, until it finally indicates a molecular weight of 64; that is, it indicates that there are two atoms of sulphur in the molecule. This shows that the complex molecules of sulphur containing six atoms are gradually broken up by heat into the simpler molecules of two atoms. Other similar cases are known. This suggests the possibility that the molecular weights of all substances may be less in the form of a gas than they are in the solid condition.

However this may be, it makes no difference in our calculations and deductions based upon the formulæ we use.

We see that our formulæ and equations, when they refer to gaseous elements and compounds, express the volume relations of these gases as well as the general relations described in Chapter V.

When marsh gas burns, the chemical change can be expressed by the following equation:—

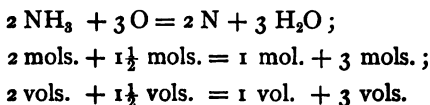


Remembering that equal volumes of gases contain an equal number of molecules, and that the molecules of the elementary gases each contain two atoms, we see that the

burning of one molecule of marsh gas requires four atoms or two molecules of oxygen, and there is formed one molecule of carbon dioxide and two molecules of water vapor ; or, one volume of marsh gas requires for burning two volumes of oxygen, and forms one volume of carbon dioxide and two volumes of water vapor. This has been proven to be true by experiment.

The same results could be obtained by calculating the weight of oxygen required to burn a liter of marsh gas, and the weight of carbon dioxide and water vapor formed, and then reducing all these weights to volumes. (Try it.)

When ammonia gas burns in oxygen, the chemical change may be expressed by the following equation, from which the volume relations can be at once deduced, knowing that oxygen and nitrogen each contains two atoms in the molecule : —

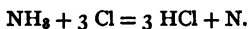


### QUESTIONS IN REVIEW.

1. What is the gas volume law of Gay-Lussac? Give an example from your own experience that illustrates this law.
2. Show how, if the theory of Avogadro is true, the knowledge that two volumes of ammonia gas give three volumes of hydrogen and one volume of nitrogen proves that the molecule of hydrogen and the molecule of nitrogen each contains two atoms.
3. What is meant by *nascent condition*? What are the phenomena of nascent condition? How are these phenomena explained?
4. What is meant by *allotropism*? How is it explained?
5. What volume of oxygen would be required to burn a liter of carbon monoxide, and what volume of carbon dioxide would be formed? The equation representing the change is,



6. What volume of chlorine would be required to decompose one liter of ammonia, and what volume of hydrochloric acid and of nitrogen would be formed? The equation representing the change is,



7. Ammonia and hydrochloric acid combine according to the equation,



In what volume relations do they combine? If the ammonium chloride,  $\text{NH}_4\text{Cl}$ , formed were broken up into its elementary gases, what volume of each would be formed provided there had been just a liter of ammonia to start with?

## CHAPTER XII.

**KINETIC THEORY OF GASES. — ENERGY OF PARTICLES PROPORTIONAL TO ABSOLUTE TEMPERATURE. — AVOGADRO'S THEORY DEDUCED. — BOYLE'S LAW DEDUCED. — LAWS OF DIFFUSION DEDUCED. — QUESTIONS IN REVIEW.**

**Kinetic Theory of Gases.** — The similar behavior of all gases toward changes in temperature and pressure, and the property that all have in common, of expanding indefinitely to fill completely and uniformly the containing vessel, very early led to attempts to form an hypothesis concerning the constitution of gases which should cause the simple relations we observe in gases to follow as a necessity.

The hypotheses first framed, however, were not altogether satisfactory ; they gradually became modified, and have been developed into what is now known as the *Kinetic Theory of Gases*, — a theory which explains the phenomena of gases in a very satisfactory way, and which has led to a very great increase in our knowledge concerning them.

The kinetic theory of gases may be simply stated as follows : —

*All gases are believed to be made up of invisible particles, (the molecules of the chemist). These particles are separated from one another by distances which, as compared with the diameter of the particles themselves, are considerable ; and are moving in straight lines in every conceivable direction at a high rate of speed.*

Now, if these assumptions are correct, it must happen that in a confined mass of gas the particles are continually colliding with one another and rebounding, their elasticity being assumed to be perfect, and also colliding with the sides of the vessel and rebounding into its interior. Thus it happens that the path of any given particle is not a straight line, but is a zigzag path made up of successive short straight lines. It will appear from this consideration that the speed of any given particle would not be constant, but that the average speed of all the particles would remain constant, assuming temperature and pressure to remain unchanged.

The pressure which the gas exerts upon the sides of the containing vessel is supposed to be due to the force of the continuous impact of the flying particles against the walls of the vessel.

The force or energy with which any given particle strikes the side of the vessel must depend, first, upon its mass, and second, upon its velocity or speed.<sup>1</sup> It is evident that the force varies directly as the mass; that is, for example, if the force of the blow struck by a body weighing 1 gram, moving at the rate of 1 m. per second, be called unity, then a body weighing 2 grams and moving at the same speed will strike a blow with twice the force; a body weighing 3 grams, and

---

<sup>1</sup> It is evident also that the direction of the motion of the particles is a factor in the force with which the particles strike the side of the vessel. Only those particles which are moving perpendicular to the surface strike the full blow that they are capable of striking.

But since the motion in any direction may be resolved into three components acting vertically to the three walls of a cubic vessel, and since a vessel of any shape may be considered as made up of an infinite number of small cubic vessels, it will be necessary for us here to consider only the case of particles moving perpendicular to the containing walls.

moving with the same speed, will strike a blow with three times the force ; etc.

Again, if a body weighing 1 gram were moving with a velocity of 2 m. per second, it would strike an obstacle with twice the force exerted by the same body moving 1 m. per second.

But in the case of the particles of gas confined in a vessel, the particle moving with twice the speed of another particle of equal weight would not only strike the sides of the vessel with twice the force at each impact, but would also strike them twice as often, so that the total force would be  $2 \times 2$ , or four times as great as that exerted by the more slowly moving particle.

A particle moving with three times the speed of another of equal weight, would strike a blow upon the walls of the vessel three times as hard and three times as often, so that the total force exerted would be nine times the force exerted by the slower moving particles of the same weight.

Or, in other words, the energy varies directly as the square of the velocity. To sum up, the energy varies directly as the mass and the square of the velocity.

Or, expressed in an equation, we have

$$\frac{E}{E'} = \frac{MS^2}{M'S'^2} \quad (1)$$

where  $E$  and  $E'$  are the energies,  $M$  and  $M'$  the masses of the particles, and  $S$  and  $S'$  the speed with which the particles move.

Since the pressure on the walls of the vessel is made up of the total energies of all the particles, it must be proportional to the energy of each particle multiplied by the number of particles. That is,

$$\frac{p}{p'} = \frac{En}{E'n'} \quad (2)$$



where  $n$  and  $n'$  represent the number of particles in each case.

**Energy Proportional to Absolute Temperature.**—Let us take a given mass of gas at  $0^\circ$  under a pressure  $p$ , and heat it up one degree; then the pressure remaining constant, the volume will be increased  $\frac{1}{273}$  of itself. According to the law of Boyle, if the pressure be increased  $\frac{1}{273}$  of itself to  $p'$ , the temperature remaining constant, the volume will be reduced to its original size. Or, in other words, if the volume be kept constant, changing the temperature from  $273^\circ$  to  $274^\circ$  (absolute temperature)—that is, increasing the temperature  $\frac{1}{273}$  of itself—has increased the pressure  $\frac{1}{273}$  of itself. The pressure then varies directly as the absolute temperature; or we have—

$$\frac{T}{T'} = \frac{p}{p'} = \frac{En}{E'n'} \quad (3)$$

But here we are dealing with the same mass of gas at two different temperatures; therefore the number of molecules  $n$  and  $n'$  is in each case the same, and we have

$$\frac{T}{T'} = \frac{E}{E'} \quad (4)$$

That is, the energy is proportional to the absolute temperature.

**Avogadro's Theory deduced from the Kinetic Theory.**—When we have equal volumes of gases under equal temperature and equal pressure, we know, since the temperatures are equal, that the energies of the particles are equal. The pressure being the sum of the energies of all the particles in each case, and being equal, it follows, since the energies are equal, that the number of particles must be equal. That is, in equation (2),

$$\frac{p}{p'} = \frac{En}{E'n'} \quad (2)$$

we have  $p = p'$  and  $E = E'$ , since the temperatures are equal; whence it follows that  $n = n'$ .

This is the theory of Avogadro.

**Boyle's Law deduced from the Kinetic Theory.**—In equal volumes of the same gas at the same temperature and under different pressures the weights must vary directly as the number of particles in the volume, and the specific gravities vary directly as the weights; or we have

$$\frac{\text{Sp. gr.}}{\text{Sp.' gr.}} = \frac{W}{W'} = \frac{n}{n'}. \quad (5)$$

We have seen above, equation (2),

$$\frac{p}{p'} = \frac{En}{E'n'}, \quad (2)$$

and from equation (4),

$$\frac{T}{T'} = \frac{E}{E'}. \quad (4)$$

Hence when the temperature remains constant, we have

$$T = T' \text{ and } E = E',$$

and equation (2) becomes

$$\frac{p}{p'} = \frac{n}{n'}. \quad (6)$$

Introducing this into equation (5), we have

$$\frac{\text{Sp. gr.}}{\text{Sp.' gr.}} = \frac{W}{W'} = \frac{n}{n'} = \frac{p}{p'}. \quad (7)$$

That is, we have the fact that the specific gravity of a gas varies directly as the pressure which it supports, deduced from the kinetic theory of gases. But this is the same as Boyle's law, which is founded upon experiment.

**Law of Diffusion deduced from Kinetic Theory.** — From equations (1) and (2) we have

$$\frac{p}{p'} = \frac{nm s^2}{n' m' s'^2} \quad (8)$$

In equal volumes of different gases under the same conditions of temperature and pressure  $p = p'$  and  $n = n'$ , whence

$$m s^2 = m' s'^2;$$

therefore

$$\frac{s^2}{s'^2} = \frac{m'}{m},$$

or

$$\frac{s}{s'} = \frac{\sqrt{m'}}{\sqrt{m}}.$$

That is, we conclude from the above reasoning that the speed of the particles of gases varies inversely as the square root of the mass of the particles. From the theory of Avogadro, the masses of the particles are proportional to the specific gravities of the gases. Hence we draw the conclusion that the velocity of the particles of gases is inversely proportional to the square root of their specific gravities. This deduction is in full accord with the phenomena of diffusion.

In Exp. 54 the water rises in the tube because the hydrogen particles, being lighter than the air particles, move much faster; the hydrogen, therefore, passes out through the pores of the plaster plug much more rapidly than the slower-moving particles of air pass in, so that a partial vacuum is caused within the tube, and the water rises.

Again, in Exp. 55, the hydrogen particles pass into the tube faster than the air particles move out, and in consequence there is an increased pressure in the tube, and some of the gas is forced out at the bottom through the water.

It has been found that gases differ very much in the rates at which they will flow through a small hole. If an apparatus is so arranged that different gases may be made to flow under a constant pressure through a pin hole in a metallic disk, and the amounts of each gas that flow through a definite time be compared with one another, these amounts are found to be inversely proportional to the square roots of their specific gravities.

Take, for example, the two gases hydrogen and oxygen : oxygen being 16 times as heavy as hydrogen, the velocities of their particles should be to each other inversely as the square roots of their specific gravities.

$$H : O :: \sqrt{16} : \sqrt{1} ;$$

$$H : O :: 4 : 1.$$

That is, the velocity of the hydrogen particle is four times as great as the velocity of the oxygen particle. It has been found by experiment that the amount of hydrogen which will flow through a small aperture in a metallic disk in a given time is just four times as great as the amount of oxygen that will flow through in the same time under the same pressure.

Simple apparatus has been designed for the determination of the specific gravity of gases by measuring the rate at which they flow through a small orifice.

Many deductions concerning gases have been reached by applying mathematics to the kinetic theory, and in so far as these deductions have been tested experimentally they have been confirmed. Without attempting to follow the theory farther, we may say in closing that it furnishes a very satisfactory explanation of the facts that all gases show a slight deviation from the laws of Boyle and Charles, and that this deviation increases as the gases approach liquefaction.

## QUESTIONS IN REVIEW.

1. What is the kinetic theory of gases?
2. Show how this theory is in accordance with Boyle's law.
3. Show how the kinetic theory of gases is in accordance with the theory of Avogadro.
4. Show how this theory is in accordance with the laws of diffusion which have been deduced by observation.
5. Calling the velocity of the hydrogen molecule unity, what are the relative velocities of the molecules of the following gases:—

- |  |       |
|--|-------|
| (a) ammonia; specific gravity is . . . . .           | 8.5   |
| (b) nitrogen; specific gravity is . . . . .          | 14.   |
| (c) hydrochloric acid; specific gravity is . . . . . | 18.25 |
| (d) carbon dioxide; specific gravity is . . . . .    | 22.   |
| (e) bromine vapor; specific gravity is . . . . .     | 80.   |

6. Given the rates of diffusion of the following gases, to determine their specific gravities, the rate of diffusion of hydrogen being taken as unity:—

- |                                      |                    |
|--------------------------------------|--------------------|
| gas (a); rate of diffusion . . . . . | 0.250 <sup>1</sup> |
| gas (b); rate of diffusion . . . . . | 0.267              |
| gas (c); rate of diffusion . . . . . | 0.183              |
| gas (d); rate of diffusion . . . . . | 0.168              |
| gas (e); rate of diffusion . . . . . | 0.148              |

<sup>1</sup> That is, if the volume of hydrogen which passes through a small aperture under a constant pressure in a given time be called one, then 0.25 will represent the volume of the gas in question, which will pass through the same aperture under the same pressure, and in the same time.

## CHAPTER XIII.

ACID. — BASE. — SALT. — ACID SALT. — BASIC SALT. — METAL. —  
NON-METAL. — CLASSIFICATION OF COMPOUNDS. — OXIDES AND  
HYDROXIDES. — CARBONATES. — CHLORIDES. — SULPHATES. —  
NITRATES. — QUESTIONS IN REVIEW.

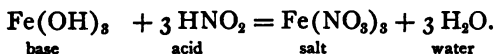
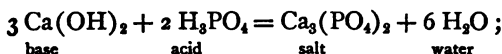
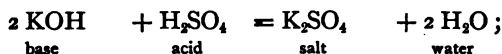
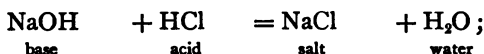
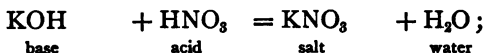
**Acids.** — Certain compounds, all of which contain hydrogen, and most of which have a sharp biting taste, and which have the power of turning blue litmus red and of changing other vegetable colors, are called *acids*. Examples are hydrochloric acid,  $\text{HCl}$ ; nitric acid,  $\text{HNO}_3$ ; sulphuric acid,  $\text{H}_2\text{SO}_4$ ; acetic acid,  $\text{HC}_2\text{H}_3\text{O}_2$ ; oxalic acid,  $\text{H}_2\text{C}_2\text{O}_4$ ; and phosphoric acid,  $\text{H}_3\text{PO}_4$ .

**Bases.** — There are other compounds called *bases*, many of which color red litmus blue and restore the vegetable colors destroyed by acids, and all of which neutralize acids. Acids and bases are opposites in character; they have the power to destroy the characteristic properties of each other. When an acid is brought in contact with a base in the proper proportions, the characteristic properties of both the acid and the base are destroyed. They are said to *neutralize* each other. They form new products which have neither acid nor basic properties, which have no action on the vegetable coloring matters, such as litmus, and which are said to be neutral. See Exps. 165–170.

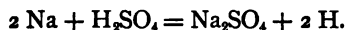
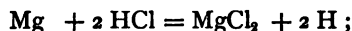
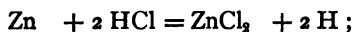
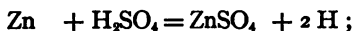
Examples of bases are potassium hydroxide,  $\text{KOH}$ ; sodium hydroxide,  $\text{NaOH}$ ; ammonium hydroxide,  $\text{NH}_4\text{OH}$ ; calcium hydroxide,  $\text{Ca}(\text{OH})_2$ ; magnesium hydroxide,

$\text{Mg}(\text{OH})_2$ ; ferric hydroxide,  $\text{Fe}(\text{OH})_3$ ; zinc hydroxide,  $\text{Zn}(\text{OH})_2$ .

**Salts.**—The substances formed when an acid neutralizes a base are called *salts*. Water is formed at the same time, as shown by the following equations:—



Salts are also formed by the action of the acids upon the metals, the metal driving the hydrogen out of the acid (see Exps. 43, 50, 51, 52, and 53), as shown in the following equations:—



**Acid Salts.**—Since the replacing power of an element is the same as its valence, there is suggested in the case of those acids which contain more than one atom of replaceable hydrogen, the possibility of a part of the hydrogen being replaced. Thus, for example, sulphuric acid,  $\text{H}_2\text{SO}_4$ , if treated with enough sodium to replace but half its hydrogen, might be expected to give such a compound as  $\text{NaHSO}_4$ , and this is just what does occur in fact. The compound

$\text{NaHSO}_4$  is a salt, and yet it is an acid, for it still has one atom of hydrogen which can be replaced by metals. Such compounds are called *acid salts*; this particular one is *acid sodium sulphate*.

It will be seen that only those acids which contain more than one atom of hydrogen can give acid salts.

Other examples of acid salts are :—

$\text{KHSO}_4$ , acid potassium sulphate ;

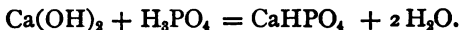
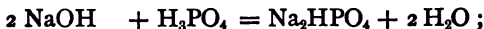
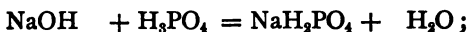
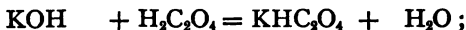
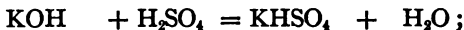
$\text{KHC}_2\text{O}_4$ , acid potassium oxalate ;

$\text{NaH}_2\text{PO}_4$ , acid sodium phosphate, or monosodium hydrogen phosphate ;

$\text{Na}_2\text{HPO}_4$ , acid sodium phosphate, or disodium hydrogen phosphate ;

$\text{CaHPO}_4$ , acid calcium phosphate.

Acid salts are usually formed whenever a base is neutralized with an excess of an acid capable of forming acid salts ; that is, with an acid containing more than one hydrogen atom.



Acids containing *one* replaceable hydrogen atom are said to be *monobasic*; those containing *two* replaceable hydrogen atoms are said to be *dibasic*; those containing *three* replaceable hydrogen atoms are *tribasic*; etc.

Only polybasic acids yield acid salts.

**Basic Salts.**—An inspection of the formulæ of bases shows us that they differ in the number of hydroxyl groups,  $(\text{OH})$ , they contain ; thus —



---

KOH	Mg(OH) <sub>2</sub>	Fe(OH) <sub>3</sub>
NaOH	Ca(OH) <sub>2</sub>	

---

They are called *monacid* bases, *diacid* bases, *triacid* bases, etc., according as they contain one, two, three, etc., hydroxyl radicals.

When the *polyacid* bases are only partially neutralized by acids, we have *basic salts*.

The following formulæ represent basic salts:—

Ca(OH)NO <sub>3</sub>	. . .	basic calcium nitrate ;
FeOHSO <sub>4</sub>	. . . . .	basic ferric sulphate ;
BiOH(NO <sub>3</sub> ) <sub>2</sub>	. . . }	basic bismuth nitrates ;
Bi(OH) <sub>2</sub> (NO <sub>3</sub> )	. . }	
FeOH(C <sub>2</sub> H <sub>3</sub> O <sub>2</sub> ) <sub>2</sub>	. }	basic ferric acetates.
Fe(OH) <sub>2</sub> (C <sub>2</sub> H <sub>3</sub> O <sub>2</sub> )	. }	

**Metals.**— Those elements which combine with oxygen, or with oxygen and hydrogen, to form compounds which have basic properties, are called *base-forming* elements or *metals*. Such elements are silver, which unites with oxygen to form the base, silver oxide, Ag<sub>2</sub>O ; potassium, which unites with hydrogen and oxygen to form the base potassium hydroxide, KOH ; magnesium, which forms the base magnesium hydroxide, Mg(OH)<sub>2</sub> ; copper, which forms the base, copper hydroxide, Cu(OH)<sub>2</sub> ; iron, which forms the base, ferric hydroxide, Fe(OH)<sub>3</sub> ; etc. The metals have the power of replacing the hydrogen in acids to form salts. They have, as a rule, a peculiar so-called “metallic” luster, and are electro-positive ; that is, when their compounds are decomposed by electricity, the metal appears at the negative pole of the battery.

**Non-Metals.**— Elements which combine with hydrogen and oxygen to form compounds with acid properties are

called *acid-forming* elements or *non-metals*. Such elements are: sulphur, in sulphuric acid,  $\text{SO}_4\text{H}_2$ ; and nitrogen, in nitric acid,  $\text{NO}_3\text{H}$ ; chlorine, in chloric acid,  $\text{ClO}_3\text{H}$ ; phosphorus, in phosphoric acid,  $\text{PO}_4\text{H}_3$ ; etc.

The non-metals do not replace the hydrogen of acids to form salts; they do not as a rule possess the "metallic" luster, they are electro-negative; that is when their compounds with the metals are decomposed by electricity, the non-metals appear at the positive pole of the battery.

All of the elements belong to one of these two classes; in most cases it is a simple matter to determine whether an element is a metal or a non-metal, but there are a few elements which under certain conditions form bases, and under certain other conditions form acids, or in other words belong now to one class and now to the other. Again, there are certain compounds which behave toward strong acids like bases, and at the same time behave toward strong bases like acids.

The distinction between the metals and the non-metals is therefore not a sharp one. Much the greater number of the elements are base-forming elements or metals.

**Classification of Compounds.** — From what has just been said we might be led to conclude that every metal yields a salt with every acid, and this, with a few exceptions, is true. In many instances, owing to the peculiar nature of the metal or of the acid, the salts are very *unstable*, tending to decompose when exposed to the air or to moisture, and consequently are prepared only with great difficulty. Such compounds are never met with in nature, and in the laboratory are formed only as the result of special processes designed for their preparation. In a few cases the compounds, if formed at all, are so unstable that chemists have not been able to obtain them.

The properties of any salt are so dependent upon both the metal and the acid from which it is formed that only by a study of the compound itself can an *accurate* knowledge of its properties be obtained.

A careful study of the elements and their compounds has shown that they may be divided into groups, such that the elements in each group are more or less closely related to one another.

No system of classification yet devised is wholly satisfactory, but the one that is, all things considered, the most satisfactory is the classification according to the periodic law. It would lead us too far to discuss here the periodic law: this subject is best taken up later, when the student has a more detailed knowledge of the elements and their compounds.

A few generalizations may be made, however, which will aid the student in gaining a knowledge of the general characteristics of certain classes of compounds.

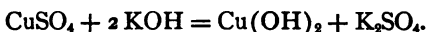
The salts of the alkali metals—lithium, sodium, potassium, rubidium, and cæsium—are all soluble in water and are all very stable; *i.e.* they are not decomposed by water nor by the atmosphere, and remain unchanged as a rule when heated. It is very difficult to get these metals free from their compounds.

The salts of the alkaline earth metals—magnesium, calcium, strontium, and barium—are many of them insoluble in water; they are only slightly less stable than the salts of the alkali metals, being undecomposed by water or by the atmosphere: many of them, however, are decomposed by high temperatures. The alkaline earth metals are very difficult to obtain free from their compounds.

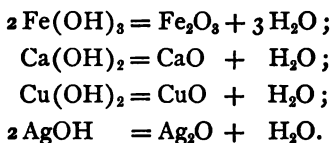
Those metals more commonly met with, like iron, lead, zinc, nickel, tin, mercury, copper, silver, gold, and plati-

num, form salts, many of which are insoluble in water and which have varying degrees of stability toward heat, but all less stable than the salts of the alkaline earth metals. All these metals may be obtained from their compounds with more or less ease, copper, silver, gold, and platinum being found in nature in the uncombined state. The noble metals, gold and platinum, are obtained whenever any of their compounds are heated.

**Oxides and Hydroxides.**—The oxides and hydroxides of all the metals except the alkali metals are insoluble or soluble with difficulty in water; hence the oxide or hydroxide is precipitated whenever a soluble hydroxide is added to a solution of any of their salts.



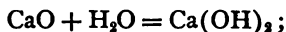
The oxide is also formed whenever any of the metals are burned in the air or in oxygen. The hydroxides of the alkali metals are not changed by heat. The hydroxides of all the other metals when heated lose water.



Some of the hydroxides, like calcium hydroxide, require a very high temperature, others need not be heated so much; copper hydroxide loses its water by boiling it with water. Others, like silver hydroxide, lose water at the ordinary temperatures, so that if they are formed at all they immediately decompose.

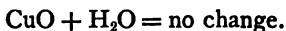
Oxides which are formed from the hydroxide only by the action of high heat, like calcium oxide  $\text{CaO}$  (ordinary

quicklime), or which cannot be formed from the hydroxide, like potassium oxide,  $K_2O$ , will unite with water very readily to form the hydroxide.



The first equation represents the familiar process of slaking lime; such changes are accompanied by the development of considerable heat.

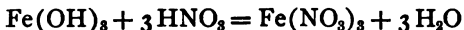
On the other hand, copper oxide does not change in water.



The oxides and hydroxides of the metals, *i.e.* the bases, are dissolved by acids with the formation of the corresponding salts of the metals, provided that the particular salt of the metal formed is soluble.

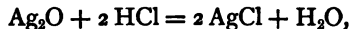
In cases where the salt formed is insoluble, the reaction is incomplete; for the hydroxide or oxide becomes covered with a coating of the insoluble salt and thus protected from the farther action of the acid.

The reaction



is complete as the ferric nitrate formed is soluble.

The reaction



on the other hand, is incomplete, because the silver chloride formed is insoluble.

**Carbonates.**—The behavior of the carbonates is very similar to that of the oxides and hydroxides.

The carbonates of all the metals, except the alkali metals, are insoluble or soluble with difficulty in water. Hence the

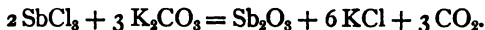
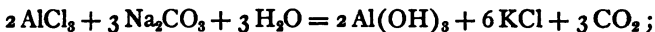
carbonate of any metal is precipitated when a soluble carbonate is added to the solution of any salt of that metal.



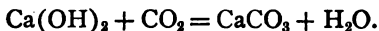
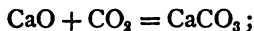
The alkaline carbonates are unchanged by heat. The carbonates of all the other metals are decomposed by heat, giving off carbon dioxide, the burning of limestone to form lime being the most familiar example of this.



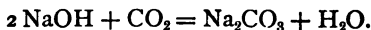
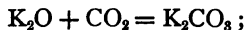
The amount of heat required to drive off the carbon dioxide varies from substances like calcium carbonate (limestone), which must be heated very hot, to substances like aluminium and antimony, the carbonates of which cannot be formed because they give off their carbon dioxide at the ordinary temperature, forming the hydroxide or oxide.



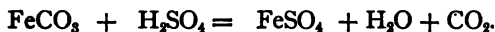
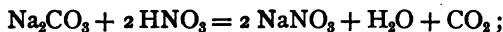
Those substances, which do not easily lose their carbon dioxide, readily take it up again after it is driven off.



The oxides and hydroxides of the alkali metals, the carbonates of which are not decomposed by heat, take up carbon dioxide to form carbonates still more readily.



All carbonates are decomposed by most acids, forming the gas carbon dioxide and the salt of the acid used.



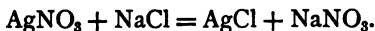
These reactions are complete only when the salt formed is soluble, for the same reasons as were given in the case of the oxides and hydroxides.

The carbonates are salts of carbonic acid,  $\text{H}_2\text{CO}_3$ ; but this acid has never been obtained in the free condition, for the reason that it breaks up into water and carbon dioxide.

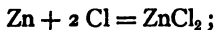


When carbon dioxide dissolves in water, the compound  $\text{H}_2\text{CO}_3$  is supposed to be in solution in the water. Water in which carbon dioxide is dissolved has an acid reaction toward litmus.

**Chlorides.**— But three metallic chlorides are insoluble in water: they are silver chloride, lead chloride, and mercurous chloride; these may be made by adding hydrochloric acid, or any soluble chloride, to a solution of the metal in question.



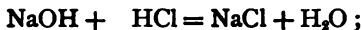
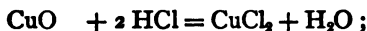
Chlorides may also be made by bringing together chlorine and the element,



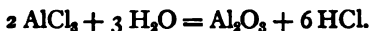
by the action of hydrochloric acid upon the metal,



or its oxide, hydroxide, or carbonate,



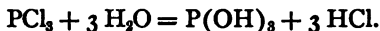
Most metallic chlorides are stable toward heat; a few, however, when heated in the presence of moisture, are decomposed.



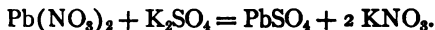
Gold and platinum chlorides when heated give off all their chlorine, and the pure metal remains.



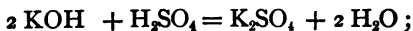
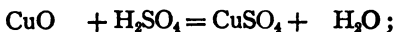
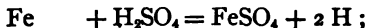
The chlorides of most of the non-metallic elements are decomposed by water.



**Sulphates.**—Most of the sulphates are soluble in water; the sulphates of calcium, strontium, barium, and lead being the insoluble sulphates, and they may be precipitated by adding sulphuric acid, or any soluble sulphate, to a solution containing the metal.

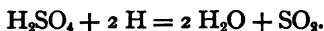


Many sulphates may be made by the action of sulphuric acid upon the metal, or upon the metallic oxide, hydroxide, or carbonate.





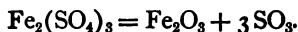
The reaction of sulphuric acid upon iron given above is typical of the action of sulphuric acid upon the metals. It sometimes happens that when the sulphuric acid is very concentrated, the nascent hydrogen formed acts upon the acid, reducing it with the formation of water and sulphur dioxide.



When copper was heated with strong sulphuric acid in Exp. 137, sulphur dioxide was obtained, and no hydrogen was given off.

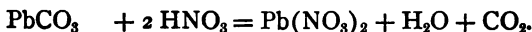
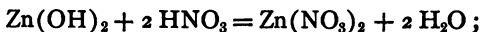
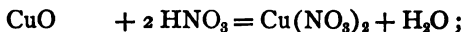


The sulphates are quite stable toward heat; the sulphates of the alkali metals and the alkaline earth metals being unchanged by heat. The sulphates of the other metals are more or less easily broken up, giving off sulphur dioxide and oxygen, and in some cases, where the heat required to decompose the sulphate is not great, sulphur trioxide is given off, at the same time the metallic oxide is formed.

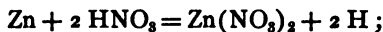


**Nitrates.** — All nitrates are soluble in water, and all nitrates are decomposed by heat.

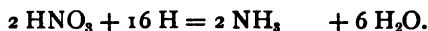
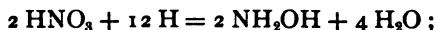
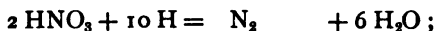
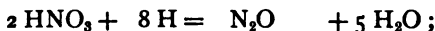
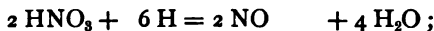
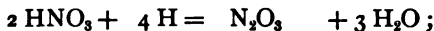
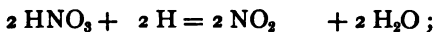
Since all nitrates are soluble, none of them can be formed by precipitation. They may be made, however, by the action of nitric acid upon the metallic oxides, hydroxides, carbonates, or upon the metals.



The action of nitric acid upon the metals is peculiar and somewhat complicated. It can be readily understood, however, if we consider the action similar to the action of the other acids upon the metals; thus, for example,

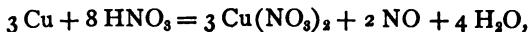


and then remember that the nascent hydrogen acts upon the nitric acid, forming a number of reduction products of nitric acid. The following reduction products have been obtained by the action of nascent hydrogen upon nitric acid:—

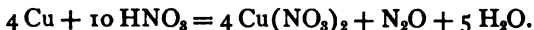


The reduction product most frequently formed in the greatest quantity by the action of nitric acid on the metals is nitric oxide, NO, although other products are formed at the same time. The products formed and the relative amounts of each depend upon the metal, the concentration of the acid, the temperature, and other physical conditions.

For the action of nitric acid upon most metals, it is customary to write the equation as if nitric oxide were the only reduction product of nitric acid formed. Such an equation, however, does not express the actual change that takes place. The action of nitric acid on copper is usually written



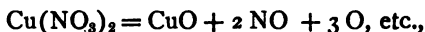
notwithstanding the fact that more or less nitrous oxide is found mingled with the nitric oxide. That is, to a certain extent the following reaction takes place :—



The nitrates of the alkali metals, when heated, lose a part of their oxygen, forming nitrites.



Other nitrates break down, forming the metallic oxide and giving off oxygen and some of the oxides of nitrogen, or even nitrogen ; thus,



several different products being formed together.

#### QUESTIONS IN REVIEW.

1. Define the terms *acid*, *base*, and *salt*.
2. Explain what is meant by *acid salt* and *basic salt*, give examples of each class, and tell how they may be prepared.
3. What is the distinction between *metals* and *non-metals* ?
4. The student should look carefully over his laboratory notes upon the oxides and hydroxides, the carbonates, chlorides, sulphates, and nitrates, and see if the observations there recorded are in accordance with the general statements made in this chapter.
5. State briefly how each class of compounds mentioned in this chapter may be prepared, and mention some of the characteristic features of each.

## CHAPTER XIV.

CHEMICAL AFFINITY. — CHEMICAL CHANGES PRODUCED BY CHEMICAL AFFINITY. — WRITING EQUATIONS. — BALANCING EQUATIONS. — QUESTIONS IN REVIEW.

**Chemical Affinity.** — The nature of the force that causes the atoms of the elements to combine with one another to form chemical compounds, is as yet unknown. It is called *chemical affinity*; it exhibits many analogies to electro-magnetic attraction, yet it is apparently different. To appreciate the importance of chemical affinity, we need but to imagine the condition of things that would result if this force ceased to act. Under such circumstances all chemical compounds would cease to exist, everything would be reduced to its chemical elements. No form of life with which we are now acquainted could exist. A vast portion of the solid matter of the earth would be resolved into invisible gases.

**Chemical Changes produced by Chemical Affinity.** — Chemical affinity holds substances together in a very intimate way, so intimate that the elements lose their characteristic properties and take on new ones. The force of chemical affinity is very great, yet we have no means of directly measuring this force. It has undoubtedly much to do with the production of the physical forces which manifest themselves as heat, light, and electricity. It is by measuring the physical force produced by a given change, which has been caused by chemical affinity, that conclusions have been reached concerning the force or strength of chemical

affinity. Thus far the most universal measure of chemical affinity used has been the heat formed. For example, chemists have judged of the affinity of chlorine for hydrogen by measuring the heat formed by the union of the two. Such measurements, however valuable they may be, do not express accurately the relative affinities of the elements for one another; for the heat measured is only the resultant of the heat produced by the union of the elements, and the heat absorbed in breaking up the molecules of the elements, and the heat produced or absorbed by the change of the physical condition of the elements during the reaction. Then, again, not all of the chemical energy is transformed into heat; investigation has shown that no chemical change takes place without the development of a certain amount of electricity.

It has been observed that in a mixture of chemicals where several possible compounds might be formed, the reaction will tend to form those compounds which develop, in their formation, the most heat. But, since the greater the chemical affinity that two elements have for one another, the greater will be the amount of heat formed by their union; this observation only confirms what we should naturally look for in such a case; namely, that the elements which have the greatest affinity for one another will tend to unite.

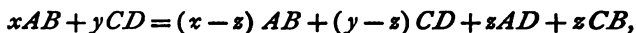
Suppose we have the two compounds  $AB$  and  $CD$ , made up of the elements  $A$ ,  $B$ ,  $C$ , and  $D$ . Now if  $A$  had an affinity for  $D$  as well as for  $B$ , and if  $C$  had an affinity for  $B$  as well as for  $D$ , when the compounds  $AB$  and  $CD$  were brought together there would tend to be an exchange; thus,



If the chemical affinity of  $A$  for  $D$  was much greater than the chemical affinity of  $A$  for  $B$ , and likewise the affinity of

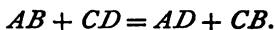
*C* for *B* much greater than the affinity of *C* for *D*, such an exchange would be nearly complete.

If, on the other hand, the affinities for the different elements were more nearly equal, there would result a state of equilibrium in which there were present the compounds *AB*, *CD*, *AD*, and *CB*; that is,



the amounts, *x*, *y*, and *z*, of each compound present depends upon the relative affinities of the elements for one another, and also upon the relative amounts of the elements present and the physical conditions. This is the simplest kind of chemical reaction.

Let us suppose, in the equation given above, that one of the products formed, *CB* for example, is insoluble; it would then be precipitated and thus thrown out of the reaction. The equilibrium of the forces which had been established by the formation and presence together of all the compounds would thus be destroyed; to reestablish it, a new amount of *CB* would be formed and precipitated. The new equilibrium thus established would be again destroyed, and the process continued until the reaction would be completed by forming all of the compound *CB* that could be formed; that is, the whole reaction would take place according to the equation

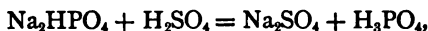


If, instead of being insoluble, and thus removed from the reaction by precipitation, *CB* were a gas and escaped from the reaction, in that way the equilibrium would be destroyed and the reaction be completed, as in the previous case.

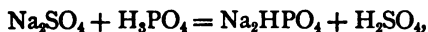
These considerations enable us better to understand the facts that have been observed; namely, that when any sub-

stances are brought together in solution, and there is a possibility of forming either an *insoluble* or a *gaseous* compound, the tendency is for that compound to be formed, and for the reaction to be complete, even though in many cases the relative affinities and the law of greatest heat formation might lead us to expect a different reaction.

Thus, at ordinary temperatures, phosphoric acid is driven out of its salts by sulphuric acid,

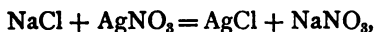


on account of the greater affinity of sulphuric acid for the sodium; but at higher temperatures the reverse reaction takes place,



because at the higher temperature the sulphuric acid becomes volatile, and is driven out of the reaction, while the phosphoric acid remains non-volatile.

Again we have this reaction taking place completely,



in spite of the fact that the affinity of chlorine for sodium is far greater than the affinity of chlorine for silver, yet the reaction is complete, because the silver chloride is insoluble and is precipitated.

The student will find many similar examples in his own experience.

Unless an insoluble or gaseous compound is formed, the reaction is never complete.

Thus it will be seen that a knowledge of the relative affinities of the elements for one another alone will not enable the chemist to say, in any given case, how they will combine. These affinities are modified and governed by so many other

conditions, that it is only through a long and careful study of the properties of an element and its compounds under various conditions, that the chemist is able to form conclusions as to how elements will unite with one another.

Indeed, chemical knowledge is not yet sufficient to enable chemists to tell with certainty what compounds will result by bringing together certain chemicals, no matter how well they may be acquainted with the properties of each, unless such an experiment has actually been performed and the products examined by some one who has made known his results. In many cases a wide knowledge of chemical reactions enables one to form a very probable conclusion as to what substances will be produced, but this conclusion must be confirmed by actual experiment before it can be put forth with confidence. In other words, our knowledge of the conditions and laws which govern the action of chemical affinity is not yet sufficient to enable us to tell just what the result will be under conditions differing from those that have been studied.

A thorough knowledge of these laws, and the ability to make use of the deductions of mathematics based upon such a knowledge, for the extension of our conceptions concerning the constitution of matter, is one of the goals of chemical thought.

The chemical elements differ very much in their chemical affinity.

We have, on the one hand, the element fluorine, which has such an affinity that for years it baffled all attempts to separate it from its compounds. This was finally accomplished by the most elaborate experiments, aided by special kinds of apparatus, and, when isolated, the most characteristic property of fluorine was its tendency to combine with other things with almost explosive violence.

On the other hand, we have the element nitrogen, which



is easily separated from its compounds, and which shows little tendency in its free condition to combine directly with other elements.

Some elements have a very wide range of affinity for other elements. Such is oxygen, which combines with all the other elements except fluorine. Other elements have a very limited range of affinity, like gold, which combines with but few elements, and then to form compounds which very readily break down, yielding the metallic gold.

In general, those elements which are chemically most unlike have the greatest tendency to combine; while those which most resemble one another chemically have least affinity for one another, in many cases forming no compound with one another. Especially is this true of the base-forming elements, which do not unite to form compounds except in very rare instances.

**Writing Equations.** — We have seen that chemical changes may be expressed in the form of equations, and that these equations express, in many cases, the actual weight relations in which the changes take place.

The student should, with the aid of the teacher, write the equations for the chemical changes with which he is dealing in the laboratory. Most of them are comparatively simple; a few of them, however, are complex and will not be understood at first.

The ability to write chemical equations correctly comes only with practice.

The student, knowing the substances started with, and at least one of the products formed, may write the equation so as to form that product, and uniting the other elements to form such compounds as he should, from his knowledge of the elements, expect to be formed. The student will here find use for all his chemical knowledge. Such equa-

tions should then be submitted to the instructor for correction. The equation can be correctly written only when some one has experimentally examined the reaction and learned all the products formed; and not always then. The student who has followed what has just been said of chemical affinity will readily understand that only in those cases where either gaseous or insoluble products are formed is the reaction complete.

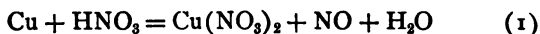
Only in such cases does the equation express the actual weight relations of the change; for only then does the equation express the whole reaction. Again, the reaction is frequently much more complex than the equation shows. Take, for example, the action of nitric acid upon the metals; it is impossible to write the equations which shall express the actual proportions in which the changes take place, for the reason that the proportion of the different reduction products of nitric acid vary with the slightest variation in the conditions of the experiment. Few reactions in inorganic chemistry are so variable as the action of nitric acid; but in organic chemistry variability is the rule, and not the exception. The equations, therefore, which are used to express the change represent but one of the numerous changes that go on. Equations, even when they express only the partial truth, aid in the understanding of the reactions, and also serve as aids to the memory. They are, therefore, much used even when they imperfectly express the true reaction.

This variability in the amount of truth expressed by different equations is frequently misleading to the student; but if he has followed the causes leading to these variations, he will soon be able to place a proper estimate upon the true value of chemical equations.

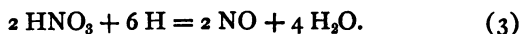
When the products started with are known, and all the

products formed are known, the balancing of the equation is a question of arithmetic only, although not always a simple question.

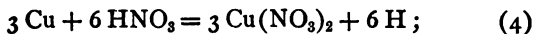
Rules have been devised for balancing equations, but it is commonly more difficult to understand the rules than it is to balance the equation. When the reaction is complex, in order to balance the equation, and also in order to more thoroughly understand the nature of the reaction, it is best to consider the changes that take place, one or two at a time. For example, we know that nitric acid acts upon copper, giving copper nitrate, nitric oxide, and water. We have, then, the equation



to be balanced. The peculiar action of nitric acid upon copper may be due to the action of nascent hydrogen upon nitric acid, which was explained in the previous chapter. If that is the case, the nascent hydrogen comes from the action of the nitric acid upon the copper; thus we have



To reduce two molecules of nitric acid to water and nitric oxide requires, according to equation (3), six atoms of nascent hydrogen; to furnish six atoms of nascent hydrogen, equation (2) must be multiplied through by three, and then becomes



combining equations (3) and (4), we get



which is equation (1) balanced.

In the study of nitric acid we found that it acts as an oxidizing agent.

In Exp. 161 it converted sulphur, in the presence of water, into sulphuric acid with the formation of nitric oxide.

We have



Before attempting to balance this equation, we must inquire how nitric acid acts as an oxidizing agent. In the previous chapter we have seen the effect of nascent hydrogen upon nitric acid; that is, we have seen how nitric acid oxidizes nascent hydrogen. It oxidizes other substances in the same way, with the formation of various reduction products of nitric acid.

If the product formed is nitric oxide, the acid breaks up in this way,



two molecules of nitric acid furnish three atoms of oxygen.

The formation of sulphuric acid from sulphur, water, and oxygen is represented thus,



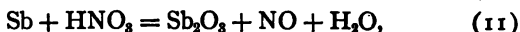
combining equations (7) and (8) we have



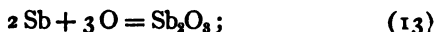
which is equation (6) balanced.

When nitric acid acts upon the metals, it produces the nitrates of all of them with but two exceptions; when it acts upon tin and antimony, it produces the oxides of these two metals. (In the case of the tin the oxide remains in combination with water,  $\text{SnO}_2 \cdot 5 \text{H}_2\text{O}$ , forming the substance known as metastannic acid.) Nitric oxide is produced at the same time.

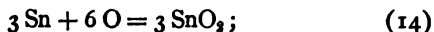
We have, therefore, the equations



to be balanced.



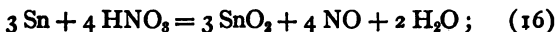
multiplying equation (12) through by 3, we get



and multiplying equation (7) through by 2, we get

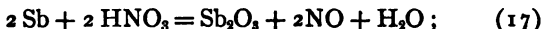


This is done so that the amount of oxygen required in equation (14) will be the same as the amount furnished in equation (15); combining equations (14) and (15), we have



which is equation (10) balanced.

Combining equations (7) and (13), we get

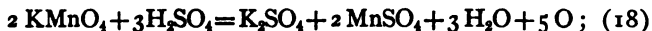


which is equation (11) balanced.

With these hints the student should be able to write the equation representing the action of nitric acid as an oxidizing agent upon any substance, provided he knows the reduction product of nitric acid, which is formed at the same time.

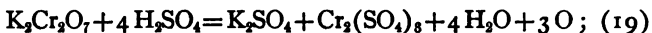
Two other substances much used as oxidizing agents in chemistry are sulphuric acid solutions of potassium permanganate,  $\text{KMnO}_4$ , and sulphuric acid solution of potassium bichromate,  $\text{K}_2\text{Cr}_2\text{O}_7$ .

Potassium permanganate acts as an oxidizing agent as follows : —



two molecules of potassium permanganate give five atoms of oxygen. All the products produced in the reaction form a colorless solution ; hence, the color of the permanganate is gradually destroyed when it acts as an oxidizing agent.

Potassium bichromate acts as an oxidizing agent as follows : —



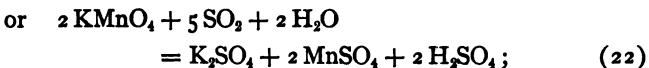
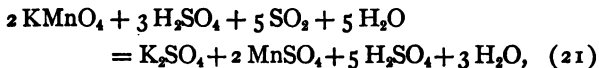
one molecule of potassium bichromate gives three atoms of oxygen.

The chromic sulphate formed is green ; hence, when potassium bichromate acts as an oxidizing agent, the red color of the bichromate is changed to the green of the chromic sulphate.

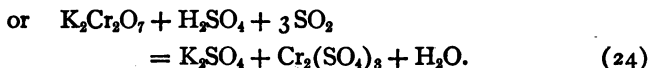
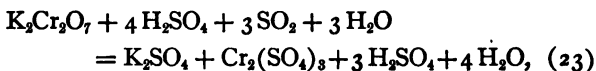
In Exp. 140, where sulphur dioxide was passed into a solution of potassium permanganate, and a solution of potassium bichromate, it was oxidized to sulphuric acid in each case ; thus,



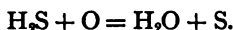
The whole reaction in the case of the permanganate is obtained by combining equations (18) and (20), and is



and in the case of bichromate, the reaction is obtained by combining equations (19) and (20).



When hydrogen sulphide is passed into the solutions of potassium permanganate and potassium bichromate, as in Exp. 144, it is oxidized and sulphur is set free ; thus,



From what has been said, the student will be able to write the equation for the whole change here.

Also to write the reactions wherever potassium bichromate or permanganate act as oxidizing agents.

By following out the hints here given, and treating other complex reactions in a similar manner, the student will not only be able to balance the equations, but will gain what is more valuable,—a clearer insight into the general character of the reactions.

### QUESTIONS IN REVIEW.

1. What is chemical affinity? Do all elements have the same chemical affinity? How do we know?

2. What other conditions besides the relative affinities of the elements influence the way in which they combine?

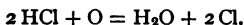
3. Explain why reactions are complete only when gaseous or insoluble products are formed.

4. When do equations express the actual weight relations between the products reacting, and when do they not?

5. Explain how it is that some equations cannot be made to express the true weight relations of the reacting products.

6. What is meant by *balancing* a chemical equation?

7. Oxidizing agents act upon hydrochloric acid to produce chlorine and water; thus



Write the equation for the action of nitric acid upon hydrochloric acid, assuming that nitric oxide is also formed.

8. Write the reactions for acid solutions of potassium permanganate,  $\text{KMnO}_4$ , and acid solutions of potassium bichromate,  $\text{K}_2\text{Cr}_2\text{O}_7$ , upon hydrochloric acid.

9. Write the reaction of nitric acid upon zinc, assuming that nitrous oxide,  $\text{N}_2\text{O}$ , is the reduction product of nitric acid formed.

10. Also write the reaction of nitric acid on zinc where ammonia,  $\text{NH}_3$ , is the reduction product of nitric acid formed.



## APPENDIX.



### THE METRIC SYSTEM OF WEIGHTS AND MEASURES.

This system, originally devised and adopted by the French, makes the *Meter* the fundamental unit. It was intended that the meter should be 0.0000001 part of the quadrant of a meridian of the earth.

With this intention an arc of the meridian, starting from the parallel of Dunkirk, in the extreme north of France, and running the entire length of France, terminating in the parallel of Barcelona in the north of Spain, was measured by Delambe and Méchain, as directed by the French government in the latter part of the last century. From this measurement the whole quadrant was computed, and the meter established as 0.0000001 part of it.

More recent measurements of the earth's surface have proved that the earth is not a perfect sphere, and that the meter really lacks about 0.0002 of an inch of being 0.0000001 of the distance from the equator to the pole.

The meter is, then, the distance measured at 0° C., between two points marked on a platinum bar, which is carefully preserved in Paris, from which copies have been taken for use.

The length of the meter in English inches is 39.3707904.

The table for linear measure is :—

10 millimeters	= 1 centimeter	= 0.3937 inches.
10 centimeters	= 1 decimeter	= 3.937 inches.
10 decimeters	= 1 meter	= 39.37 inches.
10 meters	= 1 decameter	= 32.81 feet.
10 decameters	= 1 hectometer	= 109.33 yards.
10 hectometers	= 1 kilometer	= 0.62137 miles.
10 kilometers	= 1 myriameter	= 6.2137 miles.

Fig. 42 is one decimeter divided into 10 centimeters and 100 millimeters.



FIG. 42.

The *Liter* is the unit of the measure of volume ; it is equal to one cubic decimeter or 1000 cubic centimeters.

10 milliliters	= 1 centiliter	= 0.338 fl. oz.
10 centiliters	= 1 deciliter	= 0.845 gill.
10 deciliters	= 1 liter	= { 1.0567 liquid, 0.908 dry quart.
10 liters	= 1 decaliter	= 2.6417 gallons.
10 decaleters	= 1 hectoliter	= 26.417 gallons.
10 hectoliters	= 1 kiloliter	= 264.17 gallons.
10 kiloliters	= 1 myrialiter	= 2641.7 gallons.

The unit of weight is the *Gram* ; it is the weight of a cubic centimeter or one milliliter of pure water taken at the temperature of its greatest density, 4° C., when weighed at Paris.<sup>1</sup>

<sup>1</sup> Owing to the fact that the earth is not a perfect sphere, the attraction of gravitation is not uniform upon its surface, and therefore the true weight of a body varies with its position upon the earth's surface. The weight of a body is also variable with its position, owing to the rotation of the earth about its axis. The same substance (a liter of

---

10 milligrams	= 1 centigram	= 0.15432 grains Troy.
10 centigrams	= 1 decigram	= 1.5432 grains Troy.
10 decigrams	= 1 gram	= 15.432 grains Troy.
10 grams	= 1 decagram	= 0.3527 oz. Avoirdupois.
10 decagrams	= 1 hectogram	= 3.5274 oz. Avoirdupois.
10 hectograms	= 1 kilogram	= 2.2046 lbs. Avoirdupois.
10 kilograms	= 1 myriagram	= 22.046 lbs. Avoirdupois.
10 myriagrams	= 1 quintal	= 220.46 lbs. Avoirdupois.
10 quintals	= 1 tonneau or ton	= 2204.6 lbs. Avoirdupois.

---

The metric system of weights and measures has many advantages over the systems commonly in use, among which may be mentioned : —

It is a decimal system throughout.

Its unit of volume and its unit of weight are all derived from the meter, its unit of linear measure.

---

water at 4° C., for example) weighed on a *spring* balance at the equator will weigh less than it would at the pole, for at the equator it is farther from the center of gravity of the earth. Again, it will weigh less at the equator, owing to the earth's rotation about its axis, because the centrifugal force, which at the pole entirely disappears, is greatest at the equator.

It has been calculated that the centrifugal force at the equator is  $\frac{1}{185}$  of the entire weight of a substance, and that if the earth rotated about its axis seventeen times ( $17^2 = 289$ ) as fast as it now rotates, bodies at the equator would have no weight at all. Thus we see that while the *mass* of a substance does not change, its *weight* — *i.e.* the force with which gravity acts upon its mass — varies with its position on the earth's surface. Nevertheless, the weight is used as the measure of the mass, and can be so used without error whenever the weight is determined by balances or scales in which other weights are used; for the action of gravity upon the mass of the weights used varies just the same as it does upon the mass of the substance to be weighed, so that the ratio between the weights is exactly the same as the ratio between the masses. The weight, except when determined by a *spring* balance, may be therefore considered as an exact measure of the mass, and is used with that meaning in this book.

The system is used by scientific men of all nationalities in their work and publications.

The following exercises will be helpful in gaining a familiarity with this system.

**Linear Measure.** — Purchase for yourself a metric rule, or make one as follows : —

1. Upon a piece of heavy white paper or cardboard, a little over one-half a meter long, lay off accurately a distance of one-half a meter; divide this distance into five equal parts (decimeters); then divide one of these divisions into ten equal parts (centimeters); and finally divide one of these centimeters into ten equal parts (millimeters).

I. Accustom yourself to the distances represented by 1 cm., 10, 100, 500, 1000 cm. Also estimate the length of familiar objects, and verify your estimate by actual measurement.

### Measure of Volume.

2. Take a large test tube about 15 by  $1\frac{1}{2}$  cm., and measure into this 1 c.c. from the burette, and mark the point on the tube; then measure in another centimeter, and mark that point; and so on until the tube has been filled.

3. Likewise take a large beaker and measure into it 100 c.c., marking the level of the water; then measure in another 100 c.c.; and so on until the beaker is full.

II. Use the graduated test tube and beaker to measure the volumes of liquids until you can estimate with considerable accuracy the volumes represented by 1, 10, 20, 50, 100, etc., cubic centimeters. Where exact measurements are desired, the burette should be used.

**Measure of Weight.** — For a thorough discussion of the balance, see "Elementary Physics," by Stewart and Gee, Vol. I., p. 63.

III. Before using the balance, the student should receive personal instruction as to the use and care of chemical balances.

In working with the more delicate balances, the following suggestion should always be borne in mind :—

(a) Substances (except pieces of metal or glass) must *never* be placed directly upon the scale pans, but should be weighed in appropriate vessels of glass, silver, platinum, or porcelain, *never on paper* (except where the exact weight is not desired), since this is liable to change in weight by the attraction or loss of moisture.

(b) Never weigh anything while warm.

(c) Handle the weights very carefully, using a pair of forceps, and not the fingers.

(d) Try the weights in a definite order, beginning with the largest (unless it is very evidently much heavier than the substance being weighed), then taking the next smaller, then the next, and so on.

(e) Count the weights over twice, once in the scale pan, and again as they are being put away in their places, and put down results *at once* in your notebook.

(f) Never allow the scale pans to swing violently ; always support them while weights are being put on or taken off, as well as when the substance being weighed is placed upon or removed from the pan.

(g) When through weighing, the supports of the balance beam should be raised, and the case closed.

4. Weigh a few familiar things; for example, a five-cent piece, a ten-cent piece, a twenty-five-cent piece, to become familiar with the weights represented by 1 gram, 10 grams, etc.

5. Procure a beaker capable of holding 500 c.c.; carefully weigh it upon the larger balance. Measure into it from your graduated beaker 100 c.c. of water, and weigh again; continue this process until you have in all 500 c.c. of water.

Do you get exactly the same increase in weight for each additional 100 c.c. of water? Why?

What is the relation between volume and weight in the metric system?

6. Procure a specific-gravity bottle; see that it is clean and dry; weigh it accurately upon the smaller balance, then fill it completely with distilled water; insert the stopper, and wipe off all traces of moisture from the outside, and weigh again.

How many grams of water does the bottle hold?

How many cubic centimeters?

7. Empty the bottle and fill with alcohol (or some other liquid), and weigh again.

How much alcohol does it hold?

What is the specific gravity of alcohol?

8. How many grams would 1 liter of alcohol weigh?

9. What would be the weight in grams of 1 c.m. of pure water at  $4^{\circ}$  C.?

10. The specific gravity of sulphuric acid is 1.8. How many cubic centimeters of it will weigh 1 kg.?

11. If the specific gravity of iron is 7.8, what would be the weight in kilograms of a solid cube of iron, one side of which was 10 cm.?

12. What is the relation of the units of volume and weight to the unit of linear measure in the metric system?

13. Explain the difference between *mass* and *weight*. How does difference in position on the earth's surface affect the weight of a body?

### THE THERMOMETER.

The thermometer used by scientific men differs from that used ordinarily in England and America. In the ordinary, or Fahrenheit, thermometer, the scale between the freezing point and boiling point of pure water is divided into  $180^{\circ}$ . The freezing point is  $+32^{\circ}$ , the boiling point is, therefore,  $212^{\circ}$ . Upon the Centigrade thermometer, the freezing point of water is called  $0^{\circ}$ , and the boiling point is  $100^{\circ}$ .

It is a very simple matter to convert the readings from one scale to another.

$$180^{\circ} \text{ F.} = 100^{\circ} \text{ C.};$$

$$9^{\circ} \text{ F.} = 5^{\circ} \text{ C.};$$

$$\frac{9}{5}^{\circ} \text{ F.} = 1^{\circ} \text{ C.};$$

$$1^{\circ} \text{ F.} = \frac{5}{9}^{\circ} \text{ C.}$$

From this the following formulæ are easily deduced:—

$$\frac{5}{9} (\text{F.}^{\circ} - 32) = \text{C.}^{\circ}, \text{ and } \frac{9}{5} \text{C.}^{\circ} + 32 = \text{F.}^{\circ}.$$

In Russia and Sweden a still differently graduated thermometer is in use.

In this thermometer, the Réaumur, the  $0^{\circ}$  is the freezing point of water, and  $80^{\circ}$  is the boiling point.

1. What temperature on the Centigrade scale is equal to  $0^{\circ} \text{ F.}$ ?
2. Express  $60^{\circ} \text{ C.}$  on the Fahrenheit scale.
3. The temperature of greatest density of water is  $4^{\circ} \text{ C.}$  What is the temperature expressed in the Fahrenheit scale? In the Réaumur scale?
4. Express  $60^{\circ}$  Réaumur in Centigrade and Fahrenheit degrees.
5. At what temperature would the Fahrenheit and Centigrade thermometers give the same reading? What temperature would thus be expressed on the Réaumur scale?

### WORKING WITH GLASS.

**Cutting Glass.**—Small glass tubing and rods are cut by holding them firmly upon the table, and, with a triangular file, making a small notch by drawing the file two or three times over the place where it is desired to cut the tube, then taking up the tube with both hands, one hand being each side of the notch, and the thumb pressed against the glass on the side *opposite* the notch, breaking it asunder precisely as a stick is broken.

Very large tubing, flasks, bottles, etc., are best cut by leading a crack around them.

For example, suppose it were desired to cut off the

broken top of a bottle to make a cup or dish. Procure a heavy iron wire, or rod, 4 or 5 mm. in diameter, and heat one end of it in the flame until it is red hot, then take the bottle in the left hand with the neck pointing upward, and apply the hot rod to the broken edge, and hold it there for a second or two. It is very probable that a crack will start, but if not, it may be produced with certainty by removing the heated rod, and immediately touching the spot with the moistened finger or piece of wood. (If it is desired to start a crack without beginning at the edge of the glass, make a notch with a file in the place and direction in which the crack should be, and apply the heated rod to the notch.) The crack once begun, press the red-hot end of the rod on the glass a little in front of it, when it will at once extend itself to the heated spot, and by slowly drawing the rod in the direction required, reheating it from time to time, the crack may be led, at first, a short distance downwards, and then turning at right angles horizontally round the neck so as to cut off a ring of glass including the uneven portion. It is generally not possible to carry the crack entirely around the glass until it returns into itself; a small portion of the glass will remain undivided, but the ring may be readily pulled off, a slight inequality marking the point where the crack was incomplete.

The edges of freshly cut glass tubing are extremely sharp and should be rounded off by holding the cut edge just within the flame of a Bunsen burner, a little below the top of the flame, turning it constantly; as soon as the edge is observed to be fairly rounded, it should be removed from the flame and allowed to cool.

In heating glass there are a few points that it is always well to bear in mind. Glass is a very poor conductor of heat, and when suddenly heated or cooled is apt to



crack ; pieces of glass should, therefore, be heated up gradually by bringing them into the flame slowly and removing them from the flame slowly.

Glassware which has been cooled quickly is always weaker and more apt to break during subsequent use than glassware that has been slowly cooled, or *annealed*, as it is termed.

Where the heat obtained from the Bunsen burner is not sufficient, the blast lamp should be used.

**Bordering.** — If a border is wanted on a tube, after it is cut, such as is seen upon test tubes to strengthen the tube as well as to facilitate the introduction of stoppers, it may be made by heating the edge of the tube in the flame continually, turning the tube upon its axis until the edge is soft. By then inserting into the tube a cone of charcoal and pressing upon it with continual rotation of the cone, a border is formed upon the tube, or the tube may be rapidly rotated and an iron rod or file handle held obliquely against the inner edge and the edge thus pressed out.

**Bending Tubes, Rods, etc.** — Glass tubes, when hot, may be bent in any desired shape ; for tubes not exceeding 1 cm. in diameter, the most suitable flame for heating is that of the common fish-tail or bat's wing gas burner. The tube should be held along the flame and not across it, so that the tube will be heated for 5 or 6 cm. ; turning the tube continually until it begins to bend with its own weight, it may then be bent to any desired angle. In bending tubes, avoid flattening the glass on the outer side of the bend or wrinkling it upon the inner side, as either fault weakens the glass so that it will very probably break at the bend. This is avoided by having the tube hot enough to bend without much force, and heated for a space of several centimeters, so that the bend is not all at one place.

**Choking.** — Choking, or decreasing the diameter of the tube, may be easily effected by heating the tube, continually turning it, until it is quite soft; then remove it from the flame and draw it out until it is reduced to the proper size. It can thus be drawn out to a mere thread and still remain a tube. A jet or dropping tube can be made by drawing out a tube in this manner and then cutting it off at the proper point.

**Sealing Tubes.** — Small tubes may be sealed by simply heating the end of the tube in the flame with continual rotation, when the edges will melt and run together. Larger tubes are first heated and drawn out a little as in choking. Then the flame is directed against the part of the tube which lies to one side of the most contracted portion, which is then drawn out farther until nothing but a thread of glass remains to connect the two portions. This thread of glass should then be heated just at the point where it joins the tube, when it will fuse and divide, running up into a small knob at the end of the larger tube. Direct the flame upon the small projecting knob and it will soon fuse and incorporate itself with the surrounding glass.

The whole end of the tube will, however, become thickened and contracted; it may be expanded a little by removing it from the flame and immediately forcing air very gently into it from the mouth until it takes the shape desired. If this is not accomplished at the first trial, it may be reheated and tried again. Having sealed one half of the tube successfully, the other half should be sealed in a similar way, drawing out the capillary and contracted portion of the tube. If this bit of capillary tube remaining attached after the first operation is too short or too slight to be used as a handle in drawing out the glass, it should be cut off, and while the end of the tube is heated in the

flame, a bit of waste glass held in the hand should be heated and pressed against it and the flame directed upon the point of junction; the two pieces will then adhere, and by applying the flame at the proper point the superfluous glass may be melted and drawn off attached to the waste glass. Moreover, if the remaining knob be too large to be neatly melted into the bottom of the tube, it may, when soft, be touched with a bit of heated glass and drawn off in a similar way.

**Blowing Bulbs.**—By heating the sealed end of a tube until it is quite soft, then removing from the flame and blowing gently into it, a bulb may be formed. A bulb may be blown at any desired point upon a tube by closing one end with a cork or rubber tube and pinchcock, then heating at the point the bulb is to occupy, and when hot removing from the flame.

To blow well-formed symmetrical bulbs, great care must be taken to get the tube evenly heated, and some experience is also required. For farther and more detailed instructions the student is referred to "The Methods of Glass Blowing," by W. A. Shenstone.

After trying each of the above operations upon some small pieces of scrap glass, the student should make the following pieces of apparatus:—

**Stirring Rods.**—Obtain a piece of glass rod about 5 mm. in diameter, about 45 cm. long. Cut it into three pieces 12, 15, and 18 cm. in length. Round the edges. Make hooks upon the ends of the rods by heating them about 2 cm. from one end until they bend from their own weight into the desired position.

**Hard Glass Tubes.**—Take a piece of hard glass tube about 25 cm. long and 7 or 8 mm. internal diameter; heat this in the middle, drawing out, and sealing both tubes as described above.

**Wash Bottle.** — (See wash bottle already prepared.) Procure flask with cork to fit and the necessary tubes. Soften the cork by rolling it in the cork press or under your foot upon a clean floor. Bore two parallel holes through it by means of a cork borer. These perforations should be cylindrical and of less diameter than the glass tubes they are to receive; they may be enlarged by the use of a rat-tail file. Prepare the tubes and fit them to the cork as shown in the wash bottle which you have as a pattern.

TABLE I.

COMBINING, OR ATOMIC WEIGHTS, AND SYMBOLS OF THE ELEMENTS.

Element.	Symbols.	Atomic Weight.	
		Number commonly used.	More Correct Number according to Meyer and Seubert.
ALUMINIUM . . . .	Al	27	27.04
ANTIMONY . . . . .	Sb (Stibium)	120	119.6
ARSENIC . . . . .	As	75	74.9
BARIUM . . . . .	Ba	137	136.86
BERYLLIUM . . . . .	Be	9	9.08
BISMUTH . . . . .	Bi	208	207.5
BORON . . . . .	B	11	10.9
BROMINE . . . . .	Br	80	79.76
CADMIUM . . . . .	Cd	112	111.7
Caesium . . . . .	Cs	133	132.7
CALCIUM . . . . .	Ca	40	39.91
CARBON . . . . .	C	12	11.97
Cerium . . . . .	Ce	141	141.2
CHLORINE . . . . .	Cl	35.5	35.37
CHROMIUM . . . . .	Cr	52.5	52.45
COBALT . . . . .	Co	59	58.6

TABLE I.—*Continued.*

Elements.	Symbols.	Atomic Weight.	
		Number commonly used.	More Correct Number according to Meyer and Seubert.
COPPER . . . . .	Cu (Cuprum)	63	63.18
Didymium . . . . .	Di	145	145.0
Erbium . . . . .	E	166	166.0
<i>Fluorine</i> . . . . .	F	19	19.06
Gallium . . . . .	Ga	70	69.9
Germanium . . . . .	Ge	72	72.31
GOLD . . . . .	Au (Aurum)	196	196.2
<i>HYDROGEN</i> . . . . .	H	1	1.00
Indium . . . . .	In	113	113.4
<i>IODINE</i> . . . . .	I	127	126.54
Iridium . . . . .	Ir	193	192.5
IRON . . . . .	Fe (Ferrum)	56	55.88
Lanthanum . . . . .	La	139	138.5
LEAD . . . . .	Pb (Plumbum)	207	206.39
LITHIUM . . . . .	Li	7	7.01
MAGNESIUM . . . . .	Mg	24	23.94
MANGANESE . . . . .	Mn	55	54.8
MERCURY . . . . .	Hg (Hydrargyrum)	200	199.8
MOLYBDENUM . . . . .	Mo	96	95.9
NICKEL . . . . .	Ni	59	58.6
Niobium . . . . .	Nb	94	93.7
<i>NITROGEN</i> . . . . .	N	14	14.01
Osmium . . . . .	Os	191	191.12
<i>OXYGEN</i> . . . . .	O	16	15.96
Palladium . . . . .	Pd	106	106.2
<i>PHOSPHORUS</i> . . . . .	P	31	30.96
PLATINUM . . . . .	Pt	194	194.34
POTASSIUM . . . . .	K (Kalium)	39	39.03
Rhodium . . . . .	Rh	104	104.1

TABLE I. — *Continued.*

Elements.	Symbols.	Atomic Weight.	
		Number commonly used.	More Correct Number according to Meyer and Seubert.
Rubidium . . . . .	Rb	85	85.2
Ruthenium . . . . .	Ru	103	103.5
Scandium . . . . .	Sc	44	43.97
<i>Selenium</i> . . . . .	Se	79	78.87
<i>SILICON</i> . . . . .	Si	28	28.0
SILVER . . . . .	Ag (Argentum)	108	107.66
SODIUM . . . . .	Na (Natrium)	23	22.995
STRONTIUM . . . . .	Sr	87	87.3
<i>SULPHUR</i> . . . . .	S	32	31.98
Tantalum . . . . .	Ta	182	182
<i>Tellurium</i> . . . . .	Te	127	126.7
Thallium . . . . .	Tl	204	203.7
Thorium . . . . .	Th	232	231.96
TIN . . . . .	Sn (Stannum)	118	117.35
TITANIUM . . . . .	Ti	50	50.25
Tungsten . . . . .	W (Wolfram)	184	183.6
URANIUM . . . . .	U	240	239.8
VANADIUM . . . . .	V	51	51.1
Ytterbium . . . . .	Yb	173	172.6
Yttrium . . . . .	Y	89	89.6
ZINC . . . . .	Zn	65	64.88
ZIRCONIUM . . . . .	Zr	90	90.4

In this list those elements whose names are printed in large capitals are exceedingly abundant; those printed in small capitals are not rare, while those printed in small letters are rare. The names of the non-metallic elements are printed in *italics*.

TABLE II.

TENSION OF WATER VAPOR IN MILLIMETERS OF MERCURY FOR DIFFERENT TEMPERATURES.

Temperature, Centigrade.	Tension in Millimeters of Mercury.	Temperature, Centigrade.	Tension in Millimeters of Mercury.	Temperature, Centigrade.	Tension in Millimeters of Mercury.
- 20	0.927	25	23.550	80	354.643
- 10	2.093	26	24.988	85	433.041
0	4.600	27	26.505	90	525.450
+ 5	6.534	28	28.101	95	633.778
10	9.165	29	29.782	100	760.000
11	9.792	30	31.548	105	906.41
12	10.457	31	33.405	110	1075.37
13	11.162	32	35.359	120	1491.28
14	11.908	33	37.410	130	2030.28
15	12.699	34	39.565	140	2717.63
16	13.536	35	41.827	150	3581.23
17	14.421	40	54.906	160	4651.62
18	15.357	45	71.391	170	5961.66
19	16.346	50	91.982	180	7546.39
20	17.391	55	117.478	190	9442.70
21	18.495	60	148.791	200	11688.96
22	19.699	65	186.945	220	17390.00
23	20.888	70	233.093	224.7	25 atmospheres.
24	22.184	75	288.517		





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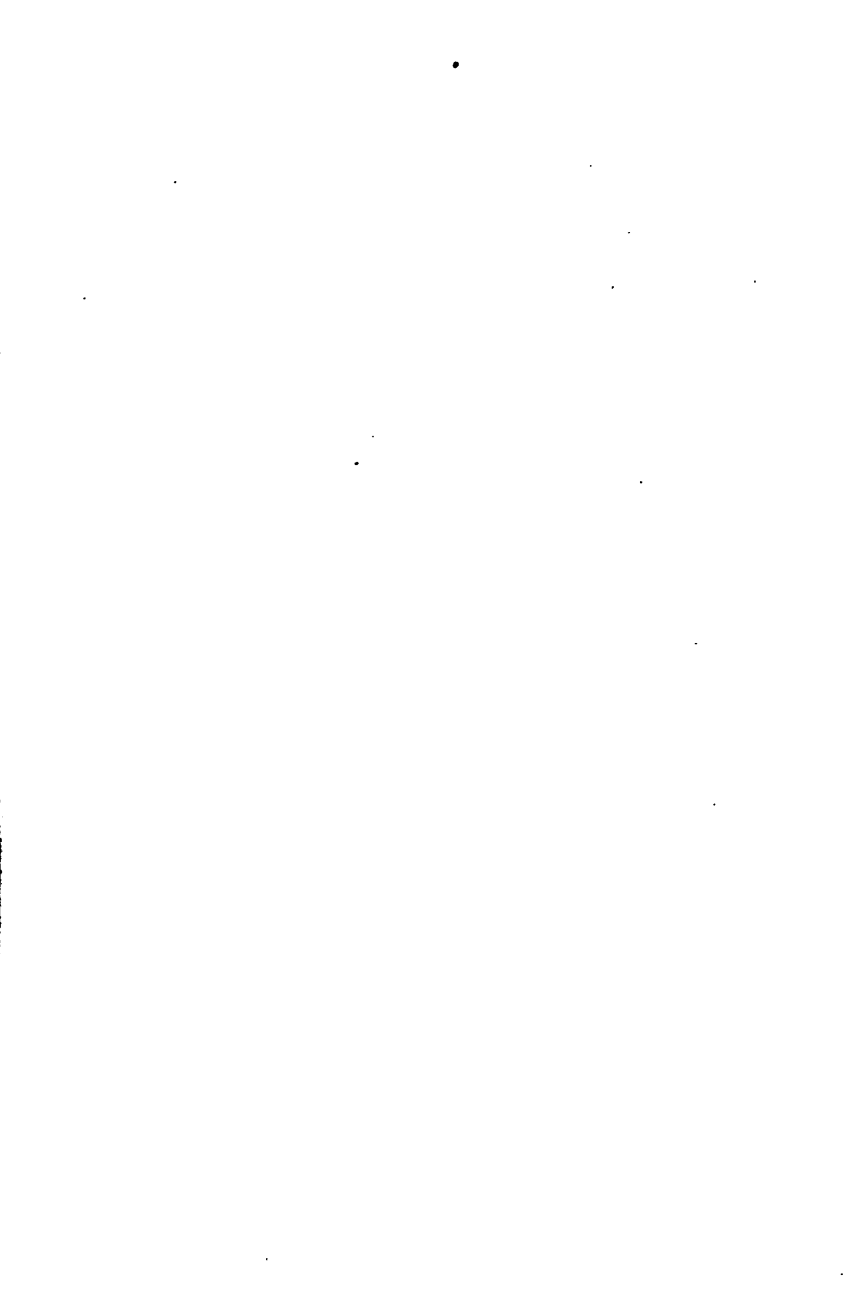
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